ISIS

INDIVIDUALIZED SCIENCE INSTRUCTIONAL SYSTEM

KITCHEN CHEMISTRY

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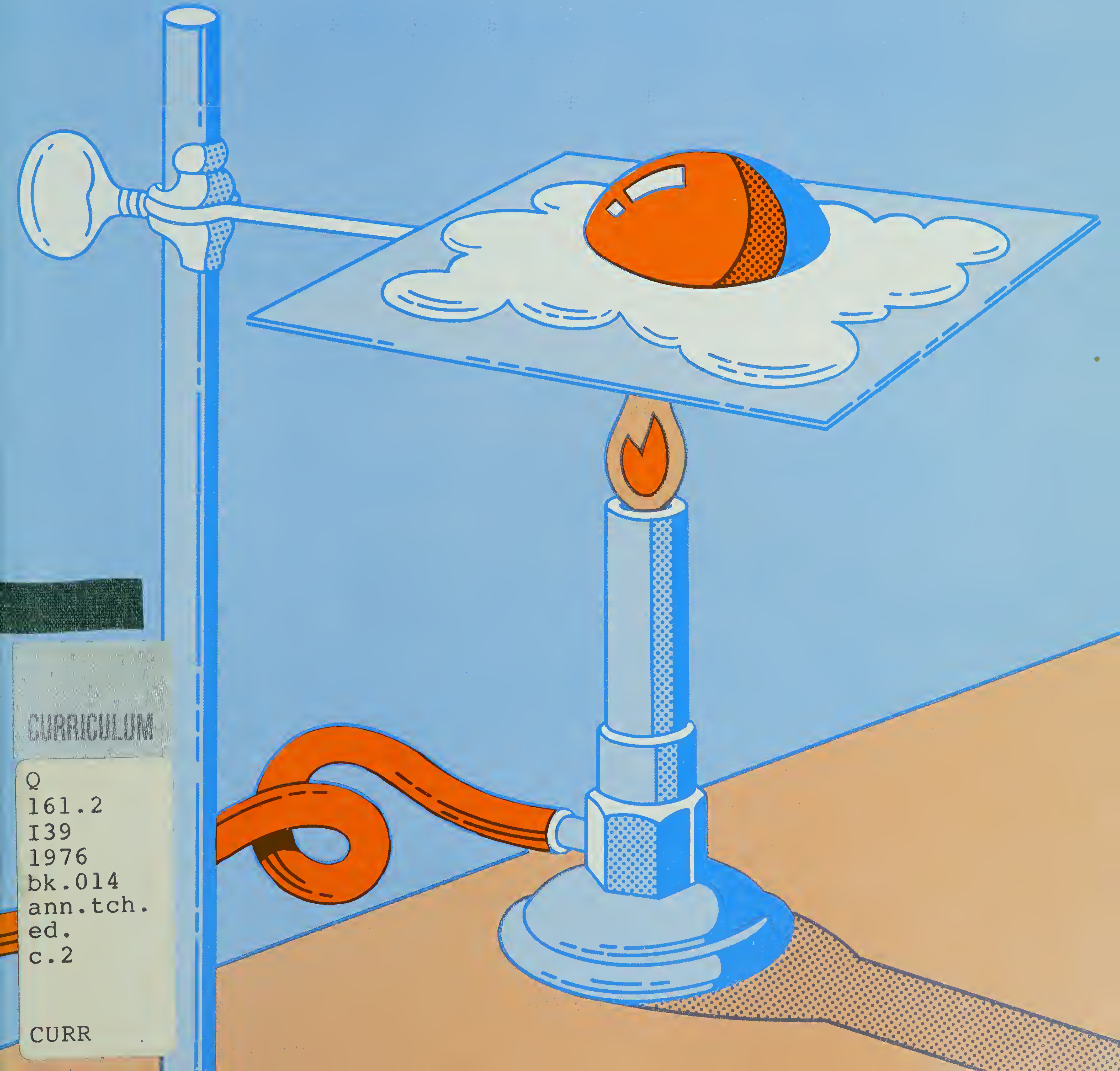

ISIS

INDIVIDUALIZED SCIENCE INSTRUCTIONAL SYSTEM

KITCHEN CHEMISTRY

ANNOTATED TEACHER'S EDITION

Ginn and Company

The work presented or reported herein was supported by a grant from the National Science Foundation. However, the opinions expressed herein do not necessarily reflect the position or policy of the National Science Foundation, and no official endorsement by that agency should be inferred.

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ANNOTATED TEACHER'S EDITION

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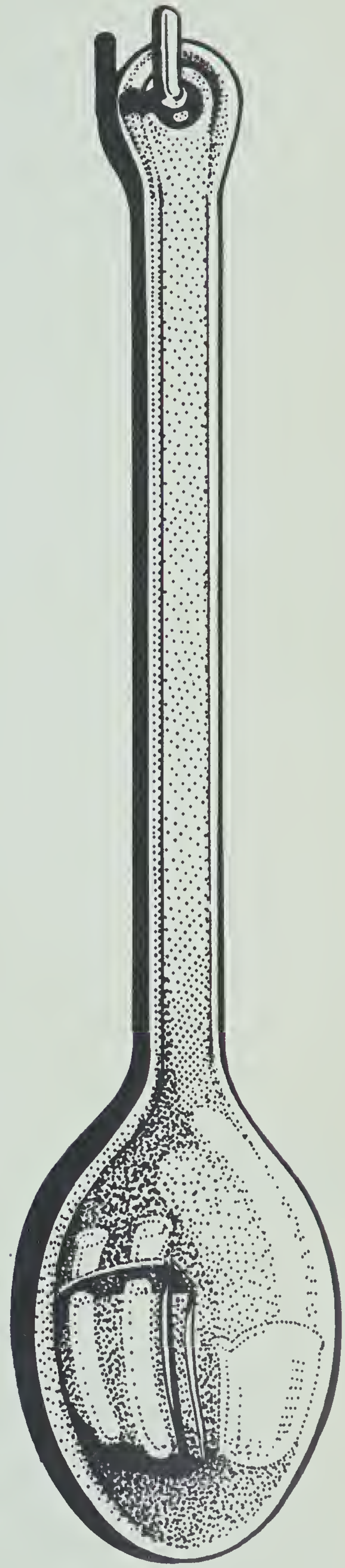

OVERVIEW

Kitchen Chemistry uses familiar kitchen substances and situations to explore solutions, suspensions, acids and bases, solubility, and crystallization.

ORGANIZATION

Kitchen Chemistry contains nine core activities, four advanced activities, and four excursion activities. The first activity in each section is a planning activity and should be done before any of the other activities in that section. Activity 9 in core should be done last. All other core activities may be done in any order.

In the core, students apply basic principles to distinguish acids and bases, look at a number of common suspensions, freeze some water, investigate the solution process, make fondant candy, learn how to control odors and tenderize meat, and learn about conservation of energy.

The advanced activities deal with the hydrogen-ion concept of acids and bases and with the freezing-point depression and the boiling-point elevation of aqueous solutions of an electrolyte (sodium chloride) and a nonelectrolyte (sugar). Students who plan to do Activities 11 and 12 should do Activity 11 first.

Students who do excursions make their own acid—base indicator from red cabbage, make some polish for leather or plastic, and make several different kinds of suspensions.

MATERIALS AND EQUIPMENT

The following tables show the quantity and the frequency of use of each item used in each activity. The activities that require no materials are not listed in the tables.

It is important to collect and organize all the materials for each minicourse before the students begin any of the activities, since the students will be working simultaneously on different activities. Having all materials readily available allows students to do the activities in the order they choose. The amount of material you will need to make available will depend on the number of lab groups that will be doing each activity. As lab groups use the "skipping option" and as they scatter themselves throughout the activities, the total amount of materials needed at one time for each activity will decrease.

CONSUMABLE ITEMS	MINIMUM MATERIALS PER LAB GROUP† PER ACTIVITY													
	2	3	4	5	6	7	8	9	11	12	13	15	16	17
*Acetic acid, $HOOCCH_3$, solution, 0.1M, in dropping bottle (drop)									2					
Alcohol, denatured, for cleanup (ml)												15		
*Alka-Seltzer (tablet)				¼										
Aluminum foil (6 cm X 6 cm piece)							1							
*Ammonia, household, in dropping bottle (ml)	1					1								1
*Ammonia concentrated solution, NH_4OH, in dropping bottle (ml)												15		
Ammonium chloride, NH_4Cl (g)										0.5				
Baking soda, $NaHCO_3$ (g)	0.1									0.5				0.1
*Barium hydroxide, $Ba(OH)_2$, solution, 0.1M, in dropping bottle (drop)									2					
Cabbage, red (g)														40
*Calcium hydroxide, $Ca(OH)_2$, saturated solution, in dropping bottle (drop)									2					
Cloth, soft												2		
Cream of tartar, $KHC_4H_4O_6$ (g)	0.1													0.1
Detergent, liquid, in dropping bottle (ml)													2	
*Fruit juice, citrus, in dropping bottle (ml)	0.1													0.1
*Gelatin, molded (small cube)							1							
Gelatin, powdered (g)													0.5	
*Hydrochloric acid, HCl, solution, 0.1M, in dropping bottle (ml)									2					
*Ice, crushed (cm^3)			500	250							500		250	
Litmus paper, blue (strip)	4								6	2				
Litmus paper, red (strip)	4								6	2				
Meat tenderizer (g)							0.5							
*Milk, sour, in dropping bottle (ml)	0.1													0.1
Oil, cooking (ml)													5	
Paper (10 cm X 10 cm piece)							1				3			
Paper towel				1	1			1					1	
Perfume or cologne (drop)						2								
*Potassium hydroxide, KOH, solution, 0.1M, in dropping bottle (drop)									2					
Potassium nitrate, KNO_3 (g)				5										
*Puffed cereal (piece)		1												
*Puffed cereal or coconut-hull fibers (test-tube full)						½								

*See "Advance Preparations."

†A *lab group* is defined as one student, a pair of students, or any size group of students that you choose.

CONSUMABLE ITEMS	MINIMUM MATERIALS PER LAB GROUP† PER ACTIVITY													
	2	3	4	5	6	7	8	9	11	12	13	15	16	17
*Shellac, 3-pound cut (ml)												6		
Sodium chloride, NaCl (g)			10	5							10			
*Sodium hydroxide, NaOH, solution, 0.1M, in dropping bottle (ml)									2					
Sodium thiosulfate, $Na_2S_2O_3 \cdot 5H_2O$ (hypo) (g)								6						
Stearic acid (octadecanoic acid), $CH_3(CH_2)_{16}COOH$ (g)												15		
Straw, soda	2									3				1
Sugar, table (g)					25						35			
Tongue depressor, wooden					1							1		
*Vinegar (acetic acid), in dropping bottle (ml)	0.1													0.1
Washing soda, Na_2CO_3 (g)										0.5				
Water, in dropping bottle (ml)						1							1	
Water, distilled (ml)										10				100
*Water, distilled, in dropping bottle (ml)	1					1		1						

*See "Advance Preparations."

†A *lab group* is defined as one student, a pair of students, or any size group of students that you choose.

NONCONSUMABLE ITEMS	MINIMUM MATERIALS PER LAB GROUP† PER ACTIVITY													
	2	3	4	5	6	7	8	9	11	12	13	15	16	17
Balance				1				1			1			1
Beaker, 100 ml					1									
Beaker, 250 ml			2	2				1				3	2	2
Brush, test tube													1	
Bunsen burner or other heat source				1	1	1		1				2	1	1
*Clamp with slotted cork					1									
Clock or watch						1		1						
Container for crushed ice, insulated			1	1							1		1	
Cup, Styrofoam			1								2			
Eraser, chalkboard		1												
Forceps or tweezers								1						
Graduated cylinder, 10 ml				1	1								1	
Grease pencil				1		1				1	1	1		1

*See "Advance Preparations."

†A *lab group* is defined as one student, a pair of students, or any size group of students that you choose.

NONCONSUMABLE ITEMS	MINIMUM MATERIALS PER LAB GROUP† PER ACTIVITY													
	2	3	4	5	6	7	8	9	11	12	13	15	16	17
Jar, baby food, with lid												1		
Knife, small, or shredder														1
Lens, hand		1											1	
*Light-beam source		1												
Medicine dropper												1		
Mortar and pestle						1								
Razor blade		1												
Ring stand, ring, and wire gauze				1	1							2		1
Ruler, metric, or metre stick						1								
Safety goggles	1		1	1	1	1	1	1	1	1	1	1	1	1
Sieve or tea strainer														1
Stirring rod									1		2			
Test tube, medium (18 mm X 150 mm)			1	2					1	3				6
*Test tube, medium (18 mm X 150 mm)						1								
Test tube, medium (18 mm X 150 mm), with stopper				2		2		1					1	
Test tube, small (13 mm X 100 mm), with stopper												2		
Test-tube holder						1		1					1	
Test-tube rack				1				1		1		1	1	1
*Thermometer, −20° C to 110° C			2		1						1			
Watch glass	1					2								
Resource Unit 5				1	1									
Resource Unit 10				1	1			1			1	1		1
Resource Unit 12						1								
Resource Unit 14			1		1									
Resource Unit 17				1	1			1						
Resource Unit 18								1						
Resource Unit 22									1					
Resource Unit 24		1			1		1	1	1					
Actions and Reactions									1	1				

*See "Advance Preparations."

†A *lab group* is defined as one student, a pair of students, or any size group of students that you choose.

ADVANCE PREPARATIONS

Activity 2

Prepare or obtain each of the following, and put them into dropping bottles.

Household ammonia. Use the nonsudsing type.

Citrus fruit juice. Use lemon juice concentrate or make a 0.1M citric acid solution (1.9 g citric acid per 100 ml of solution.)

Sour milk. Use soured whole milk, or make sour milk by adding 2 ml of 1M hydrochloric acid to 100 ml of prepared powdered milk.

Vinegar. Use household vinegar, or make 0.7M acetic acid solution by diluting 4 ml of glacial acetic acid to 100 ml.

Distilled water. Use distilled or deionized water.

Activity 3

The light-beam source could be a slide or film-strip projector or a large flashlight. A beam of sunlight coming through a shade or blind would also work.

Puffed cereal can be used in both Activities 3 and 7. Popped popcorn could also be used but does not work as well in Activity 7. The popcorn would need to be popped ahead of time.

Activity 4

This activity requires matched pairs of thermometers. Each student lab group needs two thermometers that read the same in the range of $-10°C$ to $+10°C$. Very often, vagaries of manufacture will cause any two thermometers selected at random to differ in reading by as much as 5°C, which will disturb the pedagogy. To avoid this problem, suspend several thermometers in a large beaker so that the bulbs are at the same level in the center of the beaker. Pour finely crushed ice into the beaker, covering the bulbs and almost filling the beaker. Let the assembly stand undisturbed. As some of the ice melts, the water level in the beaker will rise. When the water level covers the thermometer bulbs, check the thermometer readings. Pair up any two that have the same readings. Identify these as matched pairs with colored yarn or tape tied around the glass loop at the top of each.

The ice for the activities in this minicourse should be crushed ice or ice flakes. Ice cubes, as formed in a typical refrigerator, will not work as well. However, ice cubes can be crushed by wrapping several in a sturdy cloth and smashing them with a hammer. The crushed ice will need to be stored in a suitable container, such as a Styrofoam ice chest.

Here is another ice cream recipe, this one for strawberry ice cream. It comes from Wylie C. Senter of Doraville, Georgia.

3 packages instant vanilla pudding mix
dash of salt
835 ml light cream, such as half-and-half
250 ml of water
2 packages frozen strawberries with heavy syrup, thawed
2½ ml red food coloring

Empty the pudding mix into the freezer can. Add the dash of salt. Stir in the cream and then the water, blending them thoroughly with a spoon. Add the thawed strawberries and the food coloring. Mix everything and freeze the ice cream.

Activity 5

Quarter a few fresh Alka-Seltzer tablets ahead of time, and store them in a container with an airtight lid.

Activity 6

To prepare a slotted cork to hold the thermometer, bore a No. 5 cork and cut a slot in one side. The thermometer can be placed inside the bored hole and the whole apparatus placed in the clamp to hold it securely.

The optimum temperature for the sugar solution up to 2500 feet above sea level is 114°C. If you have a high-temperature thermometer, it may be used. A 110°C thermometer may be modified by making a file scratch or a mark four units above the 110° line. Fill the scratch with ink or paint.

At elevations of about 2500 feet above sea level, the temperature must be adjusted downward from 114°C. Use 110°C from 2500 to 4500 feet, 108°C from 4500 to 6000 feet, and 107°C above 6000 feet.

Some flavorings can be added to the fondant without harming the investigation. A few drops of mint or lemon flavorings or vanilla extract can be used. Avoid using chocolate, since it will prevent the student from seeing the action.

Activity 7

Coconut-hull fibers work best for making activated charcoal. Be sure to use the outer fibrous hull and not the hard shell. However, that may be difficult to find, since most grocery stores sell coconuts with the outer hull removed. Puffed cereal is adequate, however.

Identify a few test tubes to be used solely for the activated charcoal investigation, since they get quite coated with charcoal residue. You may want students to perform this activity in a well-ventilated area.

Household ammonia and distilled water are needed in dropping bottles; see Advance Preparations for Activity 2 (page ATE 6).

Activity 8

To make the gelatin, follow the directions on a box of unflavored or flavored dessert gelatin. Cut the gelatin into 2 to 3 cm cubes. You'll need to keep the gelatin refrigerated, or at least cooled, during class.

You might prefer having students do Activity 16, in which they make their own gelatin, before this activity.

Activity 11

Although typical dropping bottles contain about 100 ml, you may want to prepare one litre of each solution, since the weighed amounts are easier to handle. Surplus solution can be stored in large stock bottles for later use. Prepare the solutions as follows to make 1 litre of each.

Sodium hydroxide, NaOH, 0.1M. Add 4.0 g NaOH to 250 ml water and dissolve it. Dilute to 1 litre with water. Store in a plastic container.

Calcium hydroxide, $Ca(OH)_2$. Prepare a saturated solution by adding 3.0 g $Ca(OH)_2$ to 1 litre of water. Let it stand overnight. Then filter it, and save the filtrate.

Potassium hydroxide, KOH, 0.1M. Add 5.6 g KOH to 250 ml of water, and dissolve it. Dilute to 1 litre with water. Store in a plastic container.

Barium hydroxide, $Ba(OH)_2$*, 0.1M*. Prepare a solution by adding 32.0 g $Ba(OH)_2 \cdot 8\,H_2O$ to 1 litre of water.

Hydrochloric acid, HCl, 0.1M. Add 8.3 ml concentrated HCl, 12M, to 1 litre of water.

Acetic acid, $HOOCCH_3$, 0.1M. Add 5.7 ml glacial acetic acid to 1 litre of water.

CAUTION: Hydroxides and acids are corrosive and must be handled with care. Prepare the solutions in Pyrex glassware. The reactions generate heat. Always add acid to water, never the reverse.

Activity 15

Put some concentrated ammonia solution (ammonium hydroxide) in a dropping bottle. Add a caution note to the label.

Shellac can be obtained from a paint or hardware store. The 3-lb cut white shellac works best. The cut of the shellac is a measure of the density of the denatured alcohol solution. The heavier cuts may be used, but should be diluted so that they flow more easily.

BACKGROUND INFORMATION

Safety Contracts

ISIS strongly recommends that all students sign a safety contract early in the school year before beginning any minicourse. A sample safety contract can be found in *Managing ISIS*.

To help enforce the provisions of the safety contract, ISIS has placed various cautionary notes in the student's book. You will note that these cautions emphasize eye safety in particular.

Eye Safety

ISIS recommends that students wear safety goggles whenever they are working in a laboratory situation. Although a particular student may not be working with hazardous materials, other students nearby may be.

Working with Chemicals

Early in the school year, spend some time instructing your students on general laboratory safety and on appropriate precautions for working safely with chemicals. There are several general safety suggestions in *Managing ISIS*.

Disposal and Conservation of Materials

You will have to direct students about methods of safely conserving and disposing of various liquids and solids. Refer to *Managing ISIS* for general suggestions.

There are a number of activities for which you may want to give special attention to disposal and conservation. Here is one specific suggestion for Activity 9.

You can recover most of the sodium thiosulfate used in Activity 9 by having students reheat the contents of the test tube until all of the material is dissolved. Then pour the hot liquid into a heat-resistant container labeled *Used Sodium Thiosulfate*. Crystals in the jar can be obtained by filtering or decanting and drying. More crystals can be obtained from the liquid by adding denatured alcohol, which causes the sodium thiosulfate crystals to precipitate.

Measures

In the student's materials, *teaspoon* and *tablespoon* are used, when appropriate, for the volume of solids to be obtained. Such measures are used instead of asking students to weigh out solids because they are simpler and because similar-size metric spoons for dispensing small amounts are not yet readily available. The following table might be useful.

U.S. MEASURE	DRY MEASURE	LIQUID MEASURE
1 teaspoon	4.38 g	4.94 ml
1 tablespoon	14.18 g	14.78 ml
1 ounce	28.35 g	29.57 ml

Suspensions

Although information in the minicourse implies that particle size determines the stability of colloidal suspensions, other factors, such as the presence of protective colloids in the same system and the electrical charge on the colloidal particles, are also significant.

Permanent suspensions (often termed colloidal systems or dispersions) may be classified in terms of the phase being suspended and the phase doing the suspending. Figure 1 on page ATE 11 shows one such classification. There remains some disagreement, however, among various authors as to names and examples. (Note that a gas-in-gas "suspension" is a solution.)

SUSPENDED PHASE	SUSPENDING PHASE	COMMON NAME	EXAMPLES
Gas	liquid	foam (solid aerosol)	whipped cream, soap suds, beaten egg white
Gas	solid	solid foam	pumice, floating soap, marshmallow
Liquid	gas	(liquid) aerosol	fog, cloud
Liquid	liquid	emulsion	milk, mayonnaise
Liquid	solid	solid emulsion	cheese, butter
Solid	gas	solid aerosol	smoke, dust
Solid	liquid	sol	gelatin, latex paint, milk of magnesia
Solid	solid	solid sol	ruby, some alloys

Figure 1

Acid—Base Theory

The definitions of *acids* and *bases* have been very carefully developed throughout this minicourse. At core level, an acid or a base is a substance that changes the color of litmus, and acids and bases react with each other. The word *neutralize* has not been used in order to avoid confusion with *neutral* (a pH of 7).

In the advanced activities, the core definition is expanded to the hydrogen-ion concept. Based on this idea, then, a molecule or an ion that releases hydrogen ions is an acid and one that takes them is a base. Thus, sodium hydroxide is not a base, but it *contains* a base: OH^-; ammonium chloride is not an acid, but an ammonium ion is. The word *hydrolysis* is not mentioned either, but, of course, that is what is described in Activity 12 as the reaction and association with water molecules.

Baking and Carbon Dioxide

The process of baking a bread or cake is one of bubble formation and protein coagulation. The bubbles of CO_2 gas may be formed by the action of yeast on sugar (fermentation) or by the reaction of baking soda with an acid. Sour milk, buttermilk, and molasses all contain acidic substances that can react with baking soda. Baking powders are often more convenient sources of CO_2, since the reaction won't occur until the powder is moistened or heated.

Baking powders usually contain baking soda with either alum, sodium aluminum sulfate, $NaAl(SO_4)_2 \cdot 12\,H_2O$; cream of tartar, $KHC_4H_4O_6$; and sometimes tartaric acid, $H_2C_4H_4O_6$; or sodium or calcium dihydrogen phosphate, NaH_2PO_4 or $Ca(H_2PO_4)_2$. Starch or flour may be added to keep the mixture dry.

Solubility

About ninety percent of all salts dissolve to a greater extent in water at a higher temperature than at a lower one. In many cases the change is not very great, such as with sodium chloride (35.7 g/100 ml at 0°C, 39.12 g/100 ml at 100°C) in Activity 5. Others have a large increase in solubility, such as potassium nitrate (13.3 g/100 ml at 0°C, 247 g/100 ml at 100°C) in Activity 5.

Some salts, about ten percent of the total, have a decreasing solubility with increasing temperature. Some common salts of this type are lithium carbonate (1.54 g/100 ml at 0°C, 0.74 g/100 ml at 100°C), barium acetate, calcium acetate, calcium hydroxide, and calcium sulfate.

The effect of temperature on the solubility of gases is generally the opposite of the effect of temperature on the solubility of solids. Temperature is the physical measure of particle motion; an increase in temperature means that an increase in the motion of the particles has occurred. As particle motion increases, gases and solids tend toward more randomness. In a solution, gases are less random than in the vapor state. Therefore, the solubility of gases is *not* favored by an increase in temperature. Solids, on the other hand, find a solution to provide more randomness than the crystalline state. An increase in temperature then usually favors the solubility of solids.

EVALUATION SUGGESTIONS

In addition to the *Minicourse Test*, answers to which are provided with the test, you may want to use the following essay questions and laboratory performance items.

Essay Questions

Four essay suggestions are included here with model answers for your convenience. Each essay suggestion relates to material found in core activities.

1. Substances such as charcoal and baking soda can be placed in refrigerators to act as deodorizers. Although both generally have the same effect, they work differently. Explain the difference.

Answer: Charcoal acts as an adsorbent by trapping the odors. Baking soda is a base, which reacts with acid odors.

2. Besides being homogenized, milk is pasteurized — heated to kill the bacteria just before being put into containers. Why can't pasteurization be used for carbonated drinks just before they are bottled?

Answer: Heating would cause a loss of the dissolved carbon dioxide. The carbon dioxide molecules, which are responsible for the carbonation, would pick up heat energy and move faster, forming bubbles and thereby escaping from the liquid.

3. You can use a paste of baking soda and water to relieve insect bites or stings. Use the fact that baking soda is a basic substance to predict what is in a bee sting or an ant bite.

Answer: Insect bites itch and sting because the insect has injected a small amount of acid under the skin. Baking soda will react with that acid.

4. If you have a cool supersaturated solution of sugar and a cool saturated solution of sugar, which solution will form more crystals when a single sugar crystal is added to each solution? Explain why.

Answer: The supersaturated solution will form more crystals, because a supersaturated solution contains more sugar molecules than would normally dissolve at that temperature. (A saturated solution contains as many sugar molecules as can dissolve at that temperature.)

Laboratory Performance

You may want to investigate the student's ability to manipulate and use some of the equipment of the minicourse. The following tasks can help you in evaluating the student's performance.

1. Give the student strips of red and blue litmus paper and several unidentified liquids. Ask the student to identify each liquid as an acid, a base, or neither.

2. Give the student 50 ml of water and 50 g of a salt. Ask the student to make a saturated and a supersaturated solution.

REFERENCES

Benarde, M.A. *The Chemicals We Eat.* New York: McGraw-Hill Book Company, 1971.
This book, which is suitable for high school students, covers foods and food additives and their uses and safety. The chapter on "natural" foods is especially interesting.

Cobb, Vicki. *Science Experiments You Can Eat.* Philadelphia: J.B. Lippincott Company, 1972.
Thirty-eight different experiments with food, all to be done at home, are included.

Leon, Simon I. *An Encyclopedia of Candy and Ice Cream Making.* New York: Chemical Publishing Co., 1959.
Hundreds of recipes that students could try at home are included.

Peckham, Gladys C. *Foundations of Food Chemistry.* 2nd ed. Toronto, Canada: Macmillan Publishing Company, Inc., 1969.
The first chapter presents a summary of the relationship between science principles and food preparation, such as physical and chemical properties of food, solutions, and colloidal systems.

Stark, N.H. *The Formula Manual.* 3rd ed. New York: Avon Books, 1977 (paperback); Cedarburg, Wisconsin: Stark Research Corp., 1975.
With more than 500 formulas, from acne lotion to wound coating, for making useful chemical concoctions, including nineteen pages of kitchen formulas, this is a *must* reference for a practical chemistry course.

Sutton, William J. *Practical Baking.* Westport, Connecticut: Avi Publishing, 1965.
Step-by-step procedures are given for the student of baking. Numerous illustrations provide an understanding into the makeup of all types of bakery products.

ISIS

INDIVIDUALIZED SCIENCE INSTRUCTIONAL SYSTEM

KITCHEN CHEMISTRY

Ginn and Company

acknowledgments

In addition to the major effort by the ISIS permanent staff, writing conference participants, and author-consultants (listed on the inside of the back cover), the following contributed to this minicourse.

Art created by: ISIS Staff

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FOREWORD

Evidence has been mounting that something is missing from secondary science teaching. More and more, students are rejecting science courses and turning to subjects that they consider to be more practical or significant. Numerous high school science teachers have concluded that what they are now teaching is appropriate for only a limited number of their students.

As their concern has mounted, many science teachers have tried to find instructional materials that encompass more appropriate content and that allow them to work individually with students who have different needs and talents. For the most part, this search has been frustrating because presently such materials are difficult, if not impossible, to find.

The Individualized Science Instructional System (ISIS) project was organized to produce an alternative for those teachers who are dissatisfied with current secondary science textbooks. Consequently, the content of the ISIS materials is unconventional as is the individualized teaching method that is built into them. In contrast with many current science texts which aim to ''cover science,'' ISIS has tried to be selective and to limit our coverage to the topics that we judge will be most useful to today's students.

Obviously the needs and problems of individual schools and students vary widely. To accommodate the differences, ISIS decided against producing tightly structured, pre-sequenced textbooks. Instead, we are generating short, self-contained modules that cover a wide range of topics. The modules can be clustered into many types of courses, and we hope that teachers and administrators will utilize this flexibility to tailor-make curricula that are responsive to local needs and conditions.

ISIS is a cooperative effort involving many individuals and agencies. More than 75 scientists and educators have helped to generate the materials, and hundreds of teachers and thousands of students have been involved in the project's nationwide testing program. All of the ISIS endeavors have been supported by generous grants from the National Science Foundation. We hope that ISIS users will conclude that these large investments of time, money, and effort have been worthwhile.

Ernest Burkman
ISIS Project
Tallahassee, Florida

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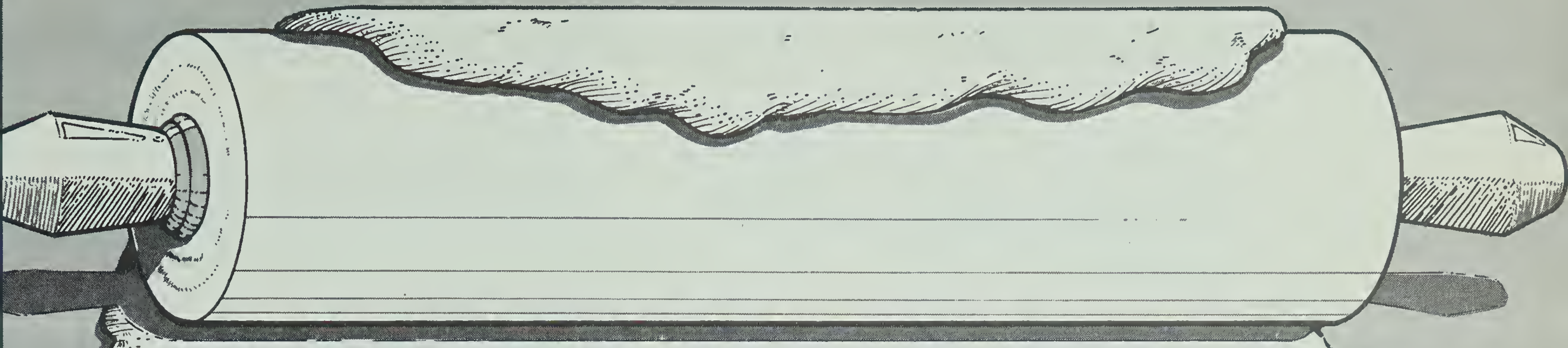


WHAT'S IT ALL ABOUT?

You may never have thought of the kitchen as a chemistry lab. But making candy or ice cream involves practical chemistry. So does using a meat tenderizer or an air freshener. You're even putting chemistry to work when you wax the floor.

Find out about these and other chemical kitchen capers in this minicourse.

ACTIVITY 1: PLANNING

If you plan to do Activity 9, do it last.

Activity 2 — Page 5

Objective 2-1: Use litmus test results to tell whether a substance is acidic, basic, or neither.

Sample Question: Suppose a substance turns blue litmus paper red. Is it acidic, basic, or neither?

Objective 2-2: Identify four substances used as acids and one used as a base in cooking.

Sample Question: Which of the following are used as acids and which as bases in cooking?
A. Citrus fruit juice
B. Sour milk
C. Vinegar
D. Cream of tartar
E. Baking soda

Objective 2-3: Describe four ways acid—base reactions are used in the kitchen.

Sample Question: Which of the following are uses of acid—base reactions in the kitchen?
A. Vinegar is mixed with corn oil to make salad dressing.
B. Lemon juice is used to remove fishy odors from hands.
C. Baking soda is mixed with cream of tartar in a cake recipe.

CORE

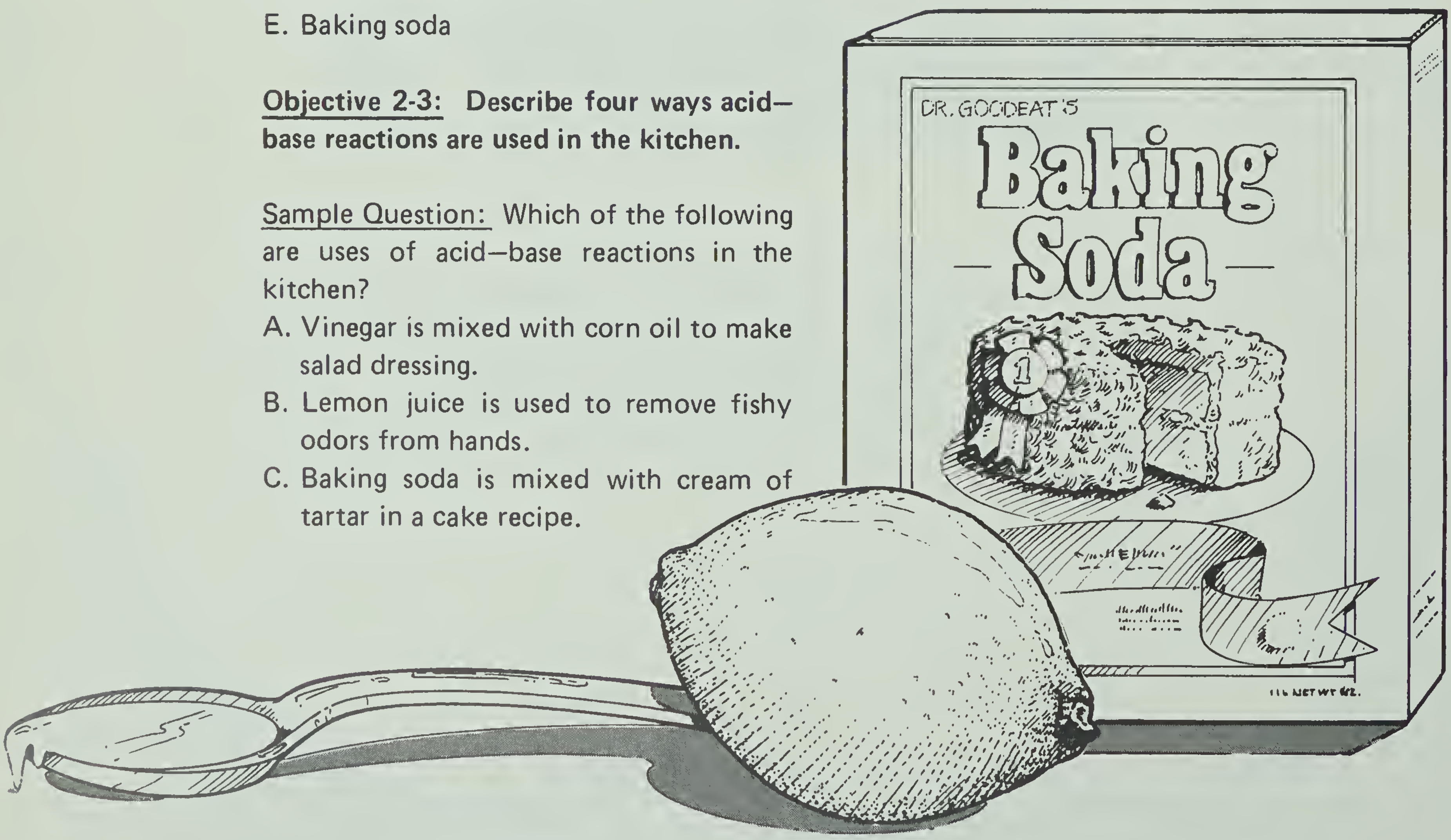
Activity 3 — Page 11

Objective 3-1: Describe temporary and permanent suspensions, and identify some examples of permanent suspensions.

Sample Question: Which of the following are permanent suspensions?
A. Milk
B. Puffed rice
C. Mayonnaise
D. Smoke

Objective 3-2: Describe how temporary and permanent suspensions differ in terms of the number of atoms in their particles.

Sample Question: In general, which has more atoms in its particles — a temporary or a permanent suspension?



Activity 4 Page 14

Objective 4-1: Explain why salt is added to the ice used in making homemade ice cream.

Sample Question: Why do you add salt to the ice in the cooling bucket when you make homemade ice cream?

A. To allow enough heat to flow from the ice cream to the ice—salt mixture
B. To make the ice—salt mixture colder than the ice cream
C. To keep the ice cream from freezing too fast
D. To make the ice cream easier to mix

Activity 5 Page 20

Objective 5-1: Tell how temperature usually affects the solubility of solids in liquids and the solubility of gases in liquids.

Sample Question: What is one reason that a warm soft drink has less carbon dioxide, CO_2, fizz than a cold soft drink?

A. A warm soft drink has less CO_2 dissolved in it than a cold soft drink does.
B. The colder a soft drink is, the less soluble CO_2 is in it.
C. CO_2 is not soluble in cold water.

Answers: 2-1. Acidic; **2-2.** Acids – A, B, C, D, base – E; **2-3.** B, C; **3-1.** A, B, C, D; **3-2.** A temporary suspension; **4-1.** A, B; **5-1.** A

Activity 6 — Page 26

Objective 6-1: Tell how saturated and supersaturated sugar solutions are made.

Sample Question: To make a supersaturated sugar solution, sugar and water are mixed and heated until all the sugar dissolves. Then

A. the sugar solution is cooled.
B. the solution is heated some more.
C. a super amount of sugar is added.

Objective 6-2: Describe how the sugar crystals in rock candy and fondant are formed.

Sample Question: If a hot sugar solution is allowed to cool slowly, without being stirred,

A. rock candy with large crystals forms.
B. smooth fondant with tiny crystals forms.
C. gritty fondant with small crystals forms.
D. no crystals form.

Activity 7 — Page 32

Objective 7-1: Tell what odors are caused by and how they travel.

Sample Question: Odors are caused by

A. all gases.
B. gases that can be seen.
C. gases that can be smelled.

Objective 7-2: Describe two methods of controlling odors.

Sample Question: Baking soda helps remove some food odors from inside a refrigerator by

A. repelling the odors.
B. overpowering the odors.
C. reacting chemically with the odors.

Activity 8 — Page 37

Objective 8-1: Describe what meat tenderizers are and what they do.

Sample Question: Meat tenderizers contain

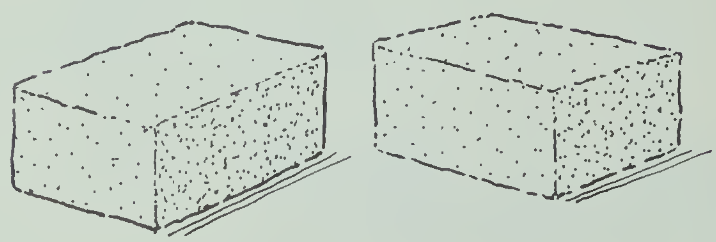
A. enzymes that keep meat from cooking too fast.
B. flavors that improve the taste of the meat.
C. enzymes that help break up protein molecules.
D. chemicals that control the odors meats give off while cooking.

Activity 9 — Page 40

Objective 9-1: Describe, in terms of the law of conservation of energy, the energy changes that occur when sodium thiosulfate crystals dissolve and then re-form from solution.

Sample Question: How does the law of conservation of energy apply to the dissolving and re-forming of sodium thiosulfate crystals?

A. Energy is neither created nor destroyed in these processes.
B. Energy is never conserved when a substance dissolves.
C. Energy is created when crystals form.



Answers: **6-1.** A; **6-2.** A; **7-1.** C; **7-2.** C; **8-1.** C; **9-1.** A

ACTIVITY 2: ACIDS AND BASES

ACTIVITY EMPHASIS: Acids and bases can be identified by using litmus papers. Certain common acid–base reactions are useful in the kitchen.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

Suppose you've just bought a hamburger and a glass of iced tea. In front of you is a white substance that's either salt or sugar. You know from experience that salt and sugar have at least one property that's very different – taste. So you'd test the white substance before using it.

Acids and bases also have very different properties. In chemistry, when you're not sure whether a substance is an acid or a base, you need to test it before using it. One way is to use specially treated paper called *litmus paper*.

You can test some common kitchen substances to find out whether they are acids or bases. You'll need the following materials.

Students may want to suggest other kitchen substances for testing. Guide them toward acidic and basic substances.

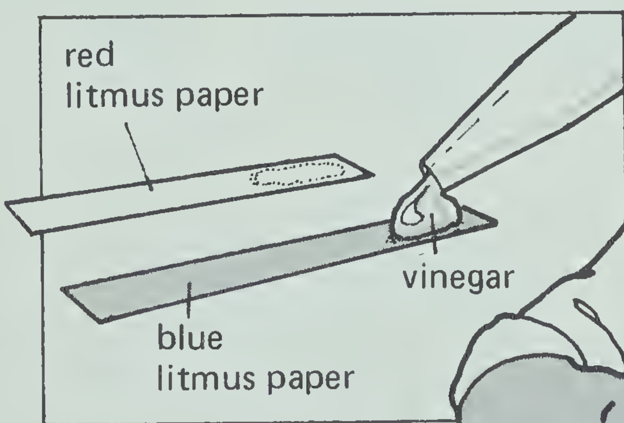
You might prefer students to use seven half-strips of litmus – one for each test.

Wood splints can be used in place of the soda straws.

- vinegar in dropping bottle
- ammonia water in dropping bottle
- distilled water in dropping bottle
- cream of tartar
- baking soda
- sour milk in dropping bottle
- citrus fruit juice in dropping bottle
- 4 strips red litmus paper
- 4 strips blue litmus paper
- 2 soda straws
- watch glass
- safety goggles

CAUTION	*Never* try a taste test on chemicals. Some chemicals are very dangerous even in small quantities.

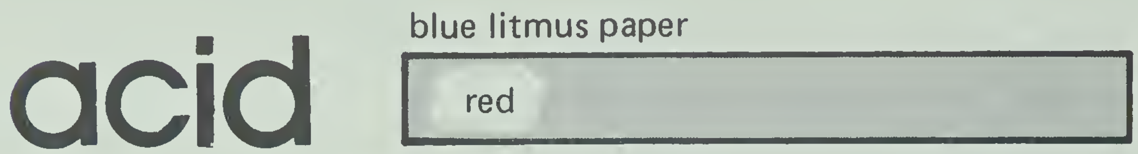
A. Get a strip of red litmus paper and a strip of blue litmus paper. Put a drop of vinegar on one end of each strip.

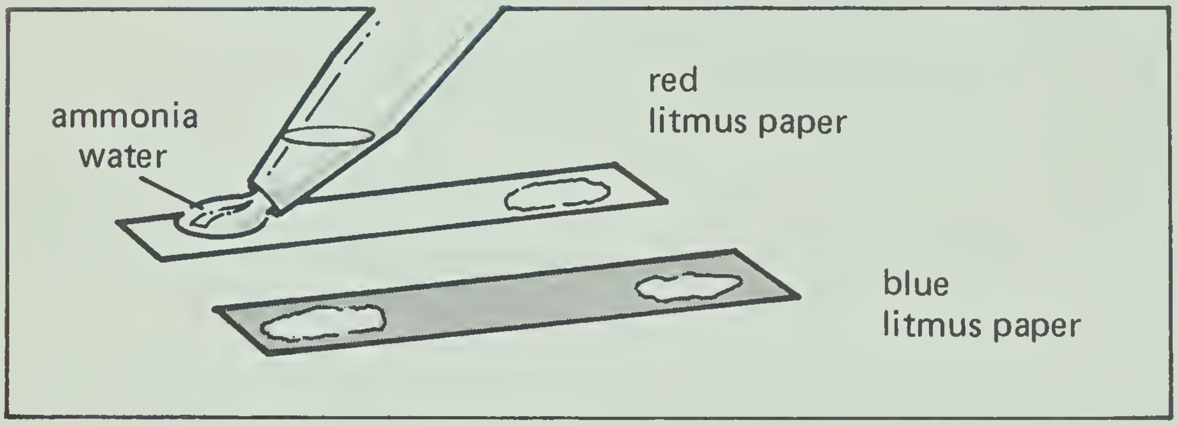


2-1. It turned red; it stayed red.

● 2-1. What happened to the color of the blue litmus strip? The red litmus strip?

Litmus is a chemical that reacts to acids and bases by changing color. Anything that turns blue litmus red is acidic. The vinegar turned the blue litmus red, so you know that it is acidic.





B. Use the same two strips of litmus paper. Put a drop of ammonia water on the unused end of each strip.

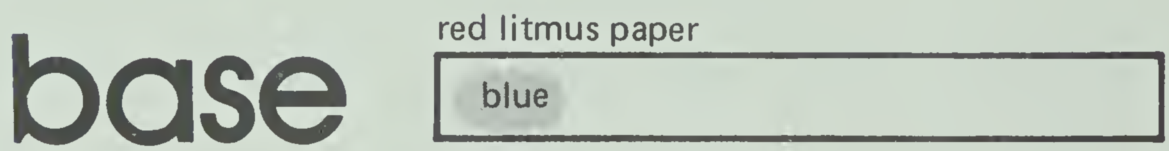
2-2. It turned blue; it stayed blue.

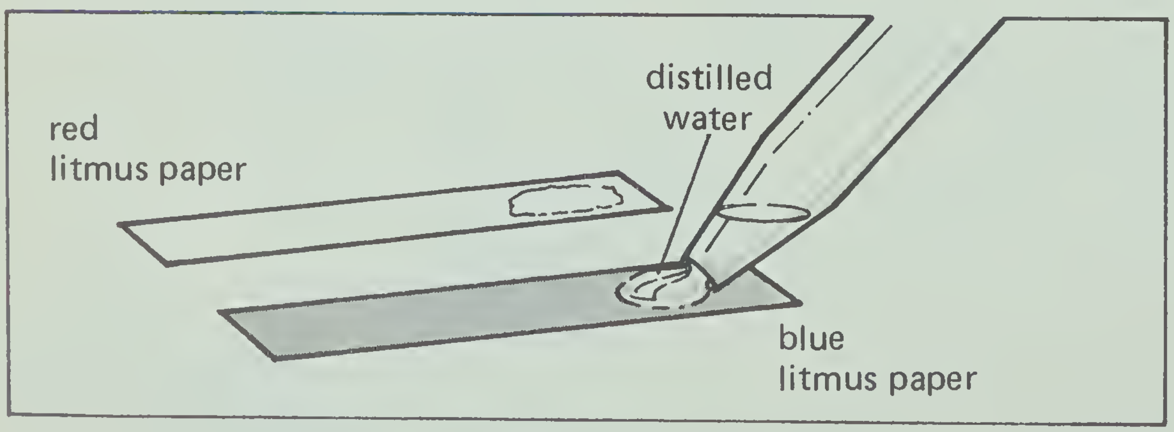
● 2-2. What happened to the color of the red litmus strip? The blue litmus strip?

2-3. No; it didn't turn the blue litmus paper red.

● 2-3. Is ammonia water acidic? Explain your answer.

Anything that makes red litmus turn blue is basic. The ammonia water turned the red litmus blue, so you know that it is basic.





C. Try distilled water next. Put a drop of distilled water on a new strip of red litmus paper. Put another drop on a new strip of blue litmus paper.

● 2-4. What happened to the color of the red litmus strip? The blue litmus strip?

2-4. They both stayed the same.

● 2-5. Is distilled water acidic or basic? Explain your answer.

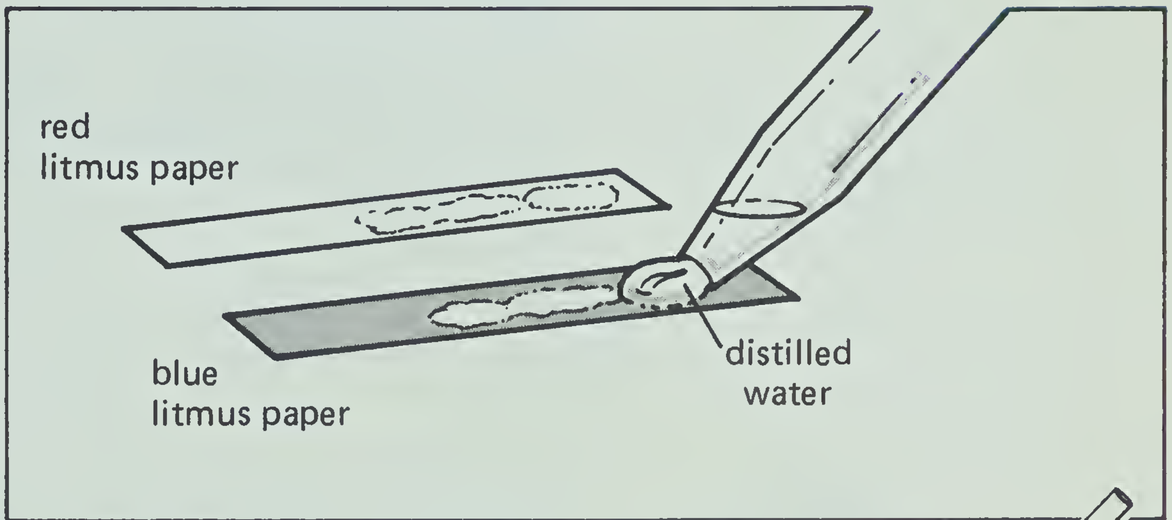
2-5. It's neither. It did not affect litmus paper.

If a substance is neither acidic nor basic, it won't change the color of litmus paper. The distilled water didn't change the the color of the litmus paper, so you know that it's neither acidic nor basic.

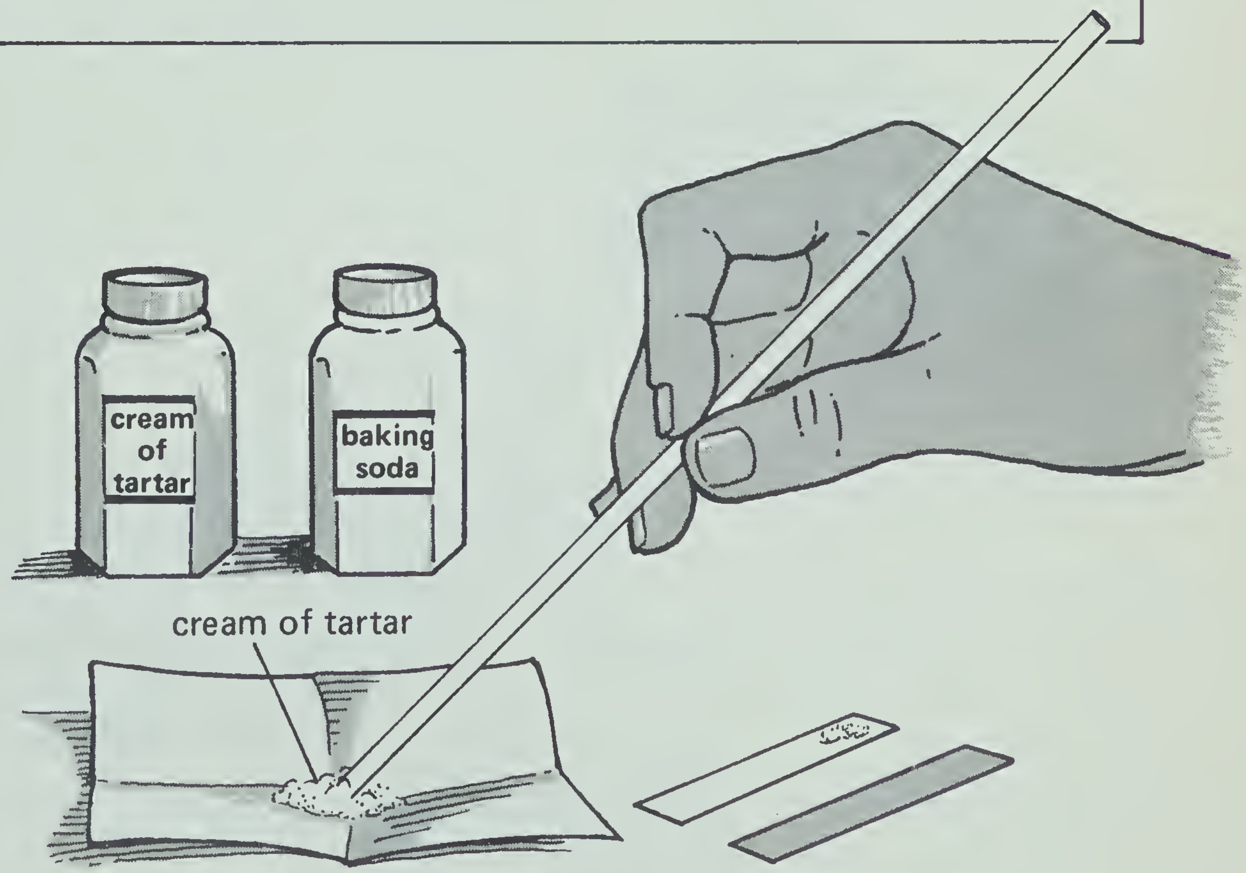
Acid :	**Blue**	**to**	**Red**
Base :	**Red**	**to**	**Blue**

Neither Acid nor Base : **No color change**

D. Test baking soda and cream of tartar next. First wet one new red strip and one new blue strip of litmus paper with several drops of distilled water.



E. Scoop up a little bit of cream of tartar in one straw and put some on one end of each strip.



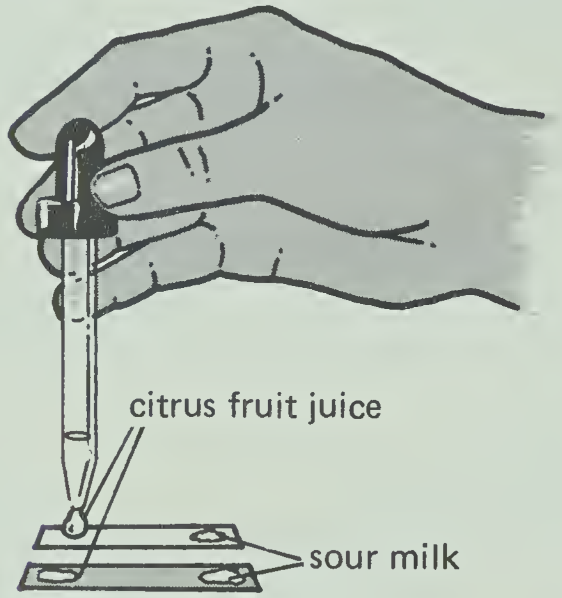
F. Scoop up a little baking soda in the other straw and put some soda on the other end of each strip.

2-6. It's an acid; it will turn blue litmus red.

☆ **2-6. Is cream of tartar an acid or a base? Explain your answer in terms of a litmus test.**

2-7. It's a base; it will turn red litmus blue.

☆ **2-7. Is baking soda an acid or a base? Explain your answer in terms of a litmus test.**



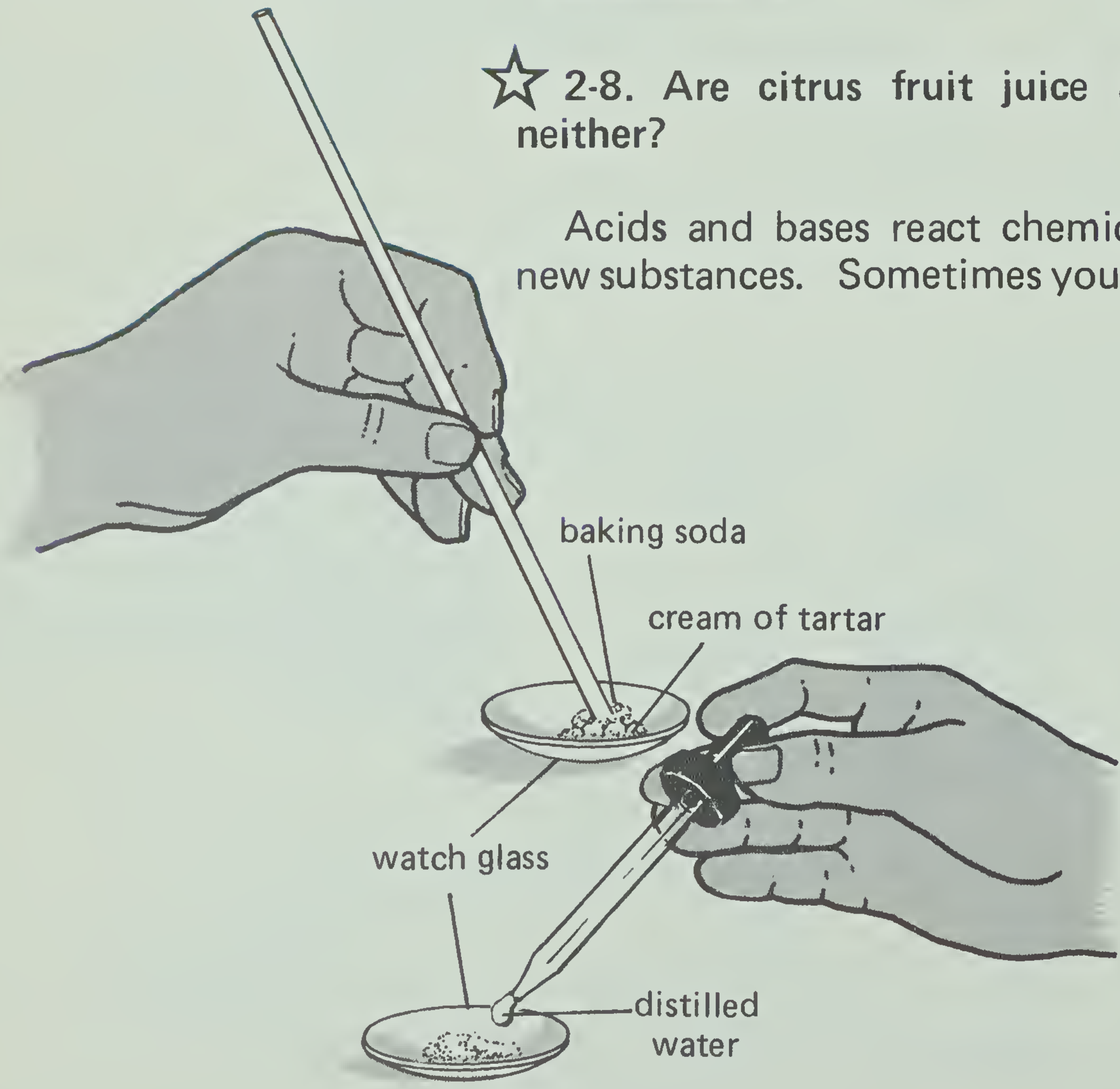
G. Try the other two liquids — sour milk and citrus fruit juice — next. Put one drop of each on a new red strip and another drop of each on a new blue strip.

2-8. Both are acidic.

The acid in sour milk is lactic acid; in citrus fruit juice the acid is citric acid. Citric acid is often added to fruit-flavored candies.

☆ **2-8. Are citrus fruit juice and sour milk acidic, basic, or neither?**

Acids and bases react chemically with each other, producing new substances. Sometimes you can see evidence of the reaction.



H. With the straw you used before for cream of tartar, put a little cream of tartar in the watch glass. With the other straw, put about the same amount of baking soda on top of the cream of tartar.

I. Mix the two powders well with either straw. Then add one or two drops of the distilled water.

● 2-9. What happened to the powders when you added the distilled water?

2-9. Bubbles were given off.

The reaction of the bicarbonate ion with an acid may be represented as follows:

$H_3O^+{}_{(aq)} + HCO_3{}^-{}_{(aq)} \rightarrow CO_{2\,(g)} + 2\,H_2O_{(l)}$

The bubbles you saw were bubbles of carbon dioxide gas, CO_2. This gas is produced whenever an acid reacts with baking soda, $NaHCO_3$.

● 2-10. Suppose you added vinegar to baking soda. What would happen? (If you aren't sure, try to find out.)

2-10. Vinegar reacts with baking soda, producing carbon dioxide.

Baking *powder* is a mixture of an acidic solid, such as cream of tartar, alum, or acid phosphate salt, with baking soda and starch or flour.

Bakers often make use of the reaction between baking soda and an acid. When baking soda and an acid are mixed in a dough or batter, tiny bubbles of CO_2 are formed. The carbon dioxide gas is what makes bread and other baked goods rise. When the mixture is baked, the bubbles expand, making the baked product porous and light.

☆ **2-11. How is an acid–base reaction used to make biscuits with sour milk and baking soda?**

2-11. The acidic sour milk reacts with the baking soda to form CO_2 bubbles in the batter. When the batter is baked, these bubbles expand and the batter rises.

When an acid and a base react, either or both of them may be used up. For example, you know from your testing that citrus fruit juices are acidic. Fish oils contain dissolved bases called *amines* [uh-MEENS]. Amines don't smell good, but most people think lemon juice – a citrus fruit juice – does. So here's a practical example of an acid–base reaction.

2-12. The lemon juice, an acid, will react with the amines in the fish oil, which are bases.

☆ **2-12. How is an acid–base reaction used to remove a fishy odor from a person's hands?**

Here's another practical use for an acid–base reaction. All aluminum pans have a natural coating of almost invisible aluminum oxide – a base. It is this coating that gets stained when certain foods are cooked in aluminum pans. You can remove the stain simply by putting something acidic, such as vinegar, on the basic aluminum oxide. Knowing some chemistry can save you a lot of hard scrubbing.

2-13. Something acidic, such as vinegar, will react with the stained aluminum oxide, which is a base.

☆ **2-13. How is an acid–base reaction used to remove stains from an aluminum pan?**

You can even use an acid–base reaction in your refrigerator. Many foods, and thus their odors, are acidic. A base such as baking soda will react with such odors and remove them.

2-14. Acidic food odors are removed by reacting with baking soda, a base.

☆ **2-14. How can an acid–base reaction be used in a refrigerator?**

ACTIVITY 3: SUSPENSIONS

ACTIVITY EMPHASIS: Temporary suspensions are temporary because they are made up of larger particles with more atoms than those in permanent suspensions. Some common permanent suspensions, such as coffee, mayonnaise, milk, and wax, are found in the kitchen.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

Something that's suspended, such as a suspension bridge, is being held up by something else. In a *chemical suspension*, one kind of substance (solid, liquid, or gas) "holds up" another kind of substance.

You can get an idea of one kind of suspension. You'll need the following materials.

chalky chalkboard eraser
light-beam source

A. Turn on the light-beam source.

B. Tap the eraser in front of the beam of light.

C. Carefully observe the chalk-dust particles in the light beam for a few minutes. Then turn off the light-beam source.

• 3-1. Describe what happened to the chalk-dust particles.

3-1. Some dust particles stayed in the air; others fell.

As you saw, some chalk-dust particles are very small, some are medium-sized, and some are very big.

• 3-2. Which particles fell first — the big ones, the medium-sized ones, or the tiny ones?

3-2. The big ones

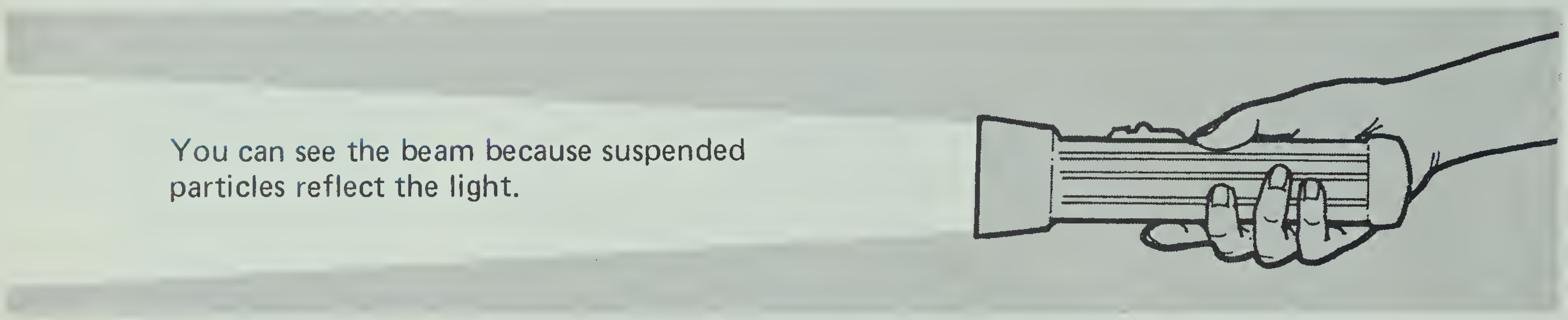
When you tapped the eraser, you made a suspension of chalk-dust particles in air. Most of the suspension was *temporary*. The bigger particles fell and were no longer in suspension. But the very smallest chalk-dust particles have probably not settled out — they've formed a *permanent* suspension in the air.

• 3-3. Which contains smaller particles, a permanent suspension or a temporary one?

3-3. A permanent suspension

Light passing through a fluid lighting up the suspended particles is called the Tyndall Effect.

You could see the chalk-dust particles in the light beam because they reflected some light back to you. The particles out of the beam were not as visible. Ordinarily, lots of very tiny particles are permanently suspended in air or water. These particles are so small that you normally don't see them. But with the right type of light, you can see millions of suspended particles in air and water.



You can see the beam because suspended particles reflect the light.

3-4. The particles reflect light back when you shine a beam of light through the suspension.

● 3-4. How is it possible to see the particles in some permanent suspensions?

3-5. A particle in a temporary suspension

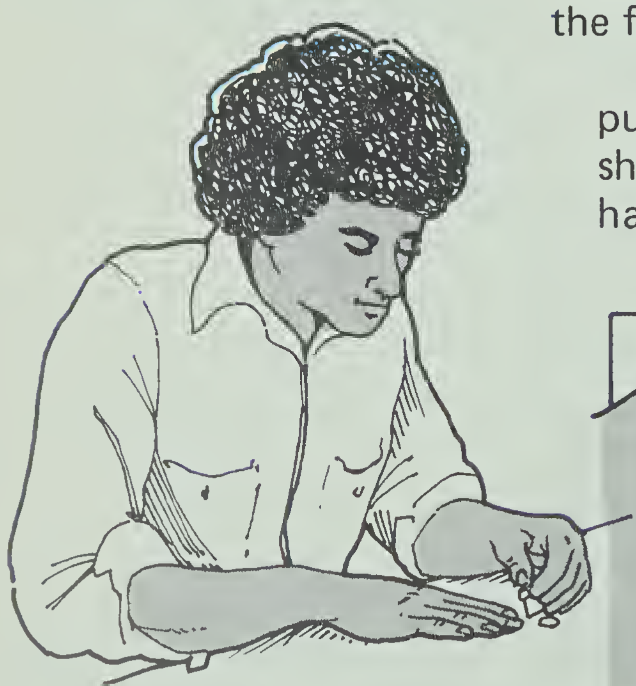
A particle in a temporary suspension usually contains millions of atoms. (If you're not sure what atoms are, see "Resource Unit 24: Atoms, Molecules, and Ions.") Even a very small permanent-suspension particle, such as a tiny water droplet in a cloud, may contain many thousands of atoms.

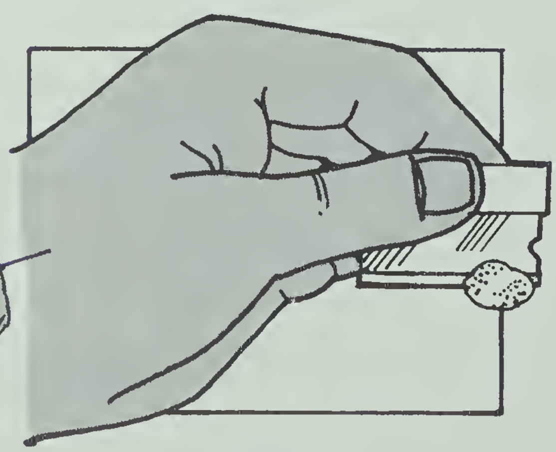
☆ **3-5. In general, which contains more atoms — a particle in a temporary suspension or one in a permanent suspension?**

Use a dissecting microscope if one is available. It shows more detail than a hand lens. You might also have students examine Styrofoam "worms" used in packing.

You've seen an example of solid particles suspended in gases. Another kind of permanent suspension is one in which gas pockets are suspended in a solid. Puffed rice and popcorn are examples. You can see this kind of suspension too. You'll need the following materials.

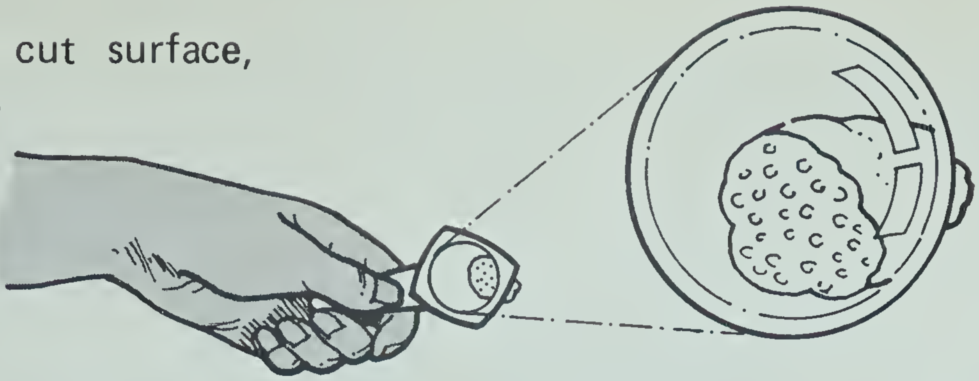
puffed cereal
sharp razor blade
hand lens





A. Cut the puffed cereal in half with the razor blade.

B. Examine the cut surface, using a hand lens.



● 3-6. What evidence did you see that puffed cereal is a suspension?

3-6. There are small holes, like little bubbles, in the cereal.

The pockets in breads and cakes that can be seen without a lens are too large to be considered in suspension.

Other examples of permanent suspensions include breads, cakes, and baked meringue on pies. Each of these contains tiny pockets of gas, some of which can be seen only with a magnifier.

You can find lots of other permanent suspensions in the kitchen. Figure 3-1 below shows some of them.

PERMANENT SUSPENSIONS

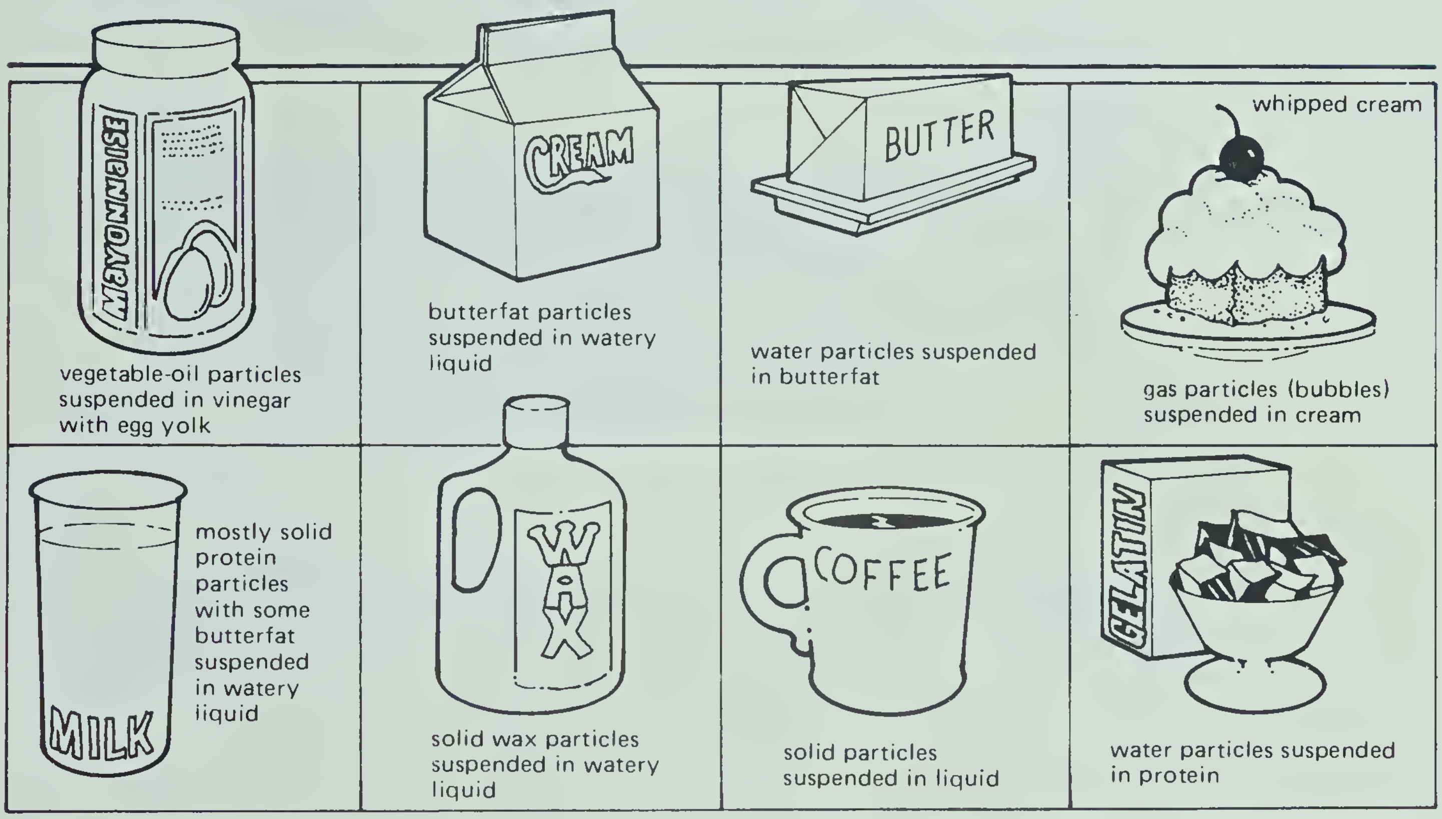


Figure 3-1

☆ **3-7. What are three ways in which temporary and permanent suspensions are different?**

3-7. The particles in temporary suspensions settle out. Those in permanent suspensions remain suspended. Permanent-suspension particles are smaller than temporary-suspension particles. Permanent-suspension particles usually have fewer atoms per particle.

☆ **3-8. List at least five permanent suspensions found in the kitchen.**

3-8. All the substances shown in Figure 3-1 (above) and popcorn, puffed rice, puffed wheat, meringue, breads, and cakes are examples.

If you'd like to make some suspensions, do "Activity 16: More on Suspensions."

ACTIVITY EMPHASIS: Heat flows from areas of higher temperature to areas of lower temperature. Salt added to an ice–water mixture lowers the temperature of the mixture, enabling the mixture to freeze ice cream.

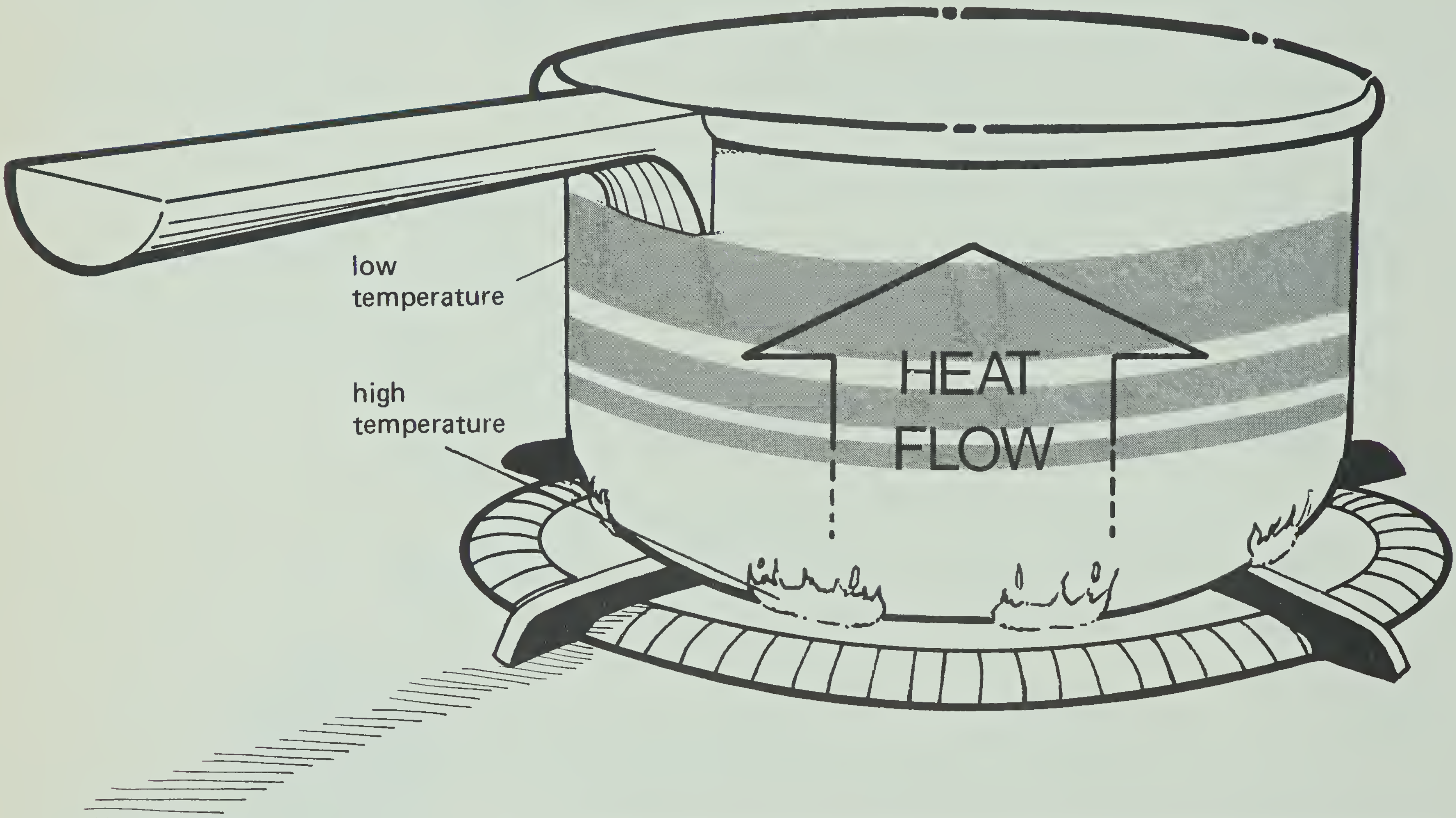
MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

ACTIVITY 4: FREEZING COLD

At the end of this activity you'll find a recipe for ice cream. You might enjoy making the ice cream and eating it. One very important step in the recipe is freezing the ice cream. To understand freezing, you need to know something about how heat moves into and out of substances.

To measure the level of heat, you use a thermometer. Thermometers measure what is called *temperature*. The scale on the thermometer gives the temperature in degrees.

Suppose you want to make some gelatin dessert. First, you heat some water on a stove. The water gets hot because heat is transferred from the burner to the pot and from the pot to the water.



4-1. It increases.

- 4-1. What happens to the temperature of water as it is heated?

When the water boils, you add it to the gelatin granules in a bowl and stir. Then you let the mixture cool. The mixture is at a higher temperature than the air in the room. Heat leaves the mixture and enters the cooler air surrounding it.

Finally, you set the gelatin by making it very cold – by putting it into the refrigerator. The mixture is at a higher temperature than the air in the refrigerator. Some more heat leaves the mixture and enters the colder air surrounding it.

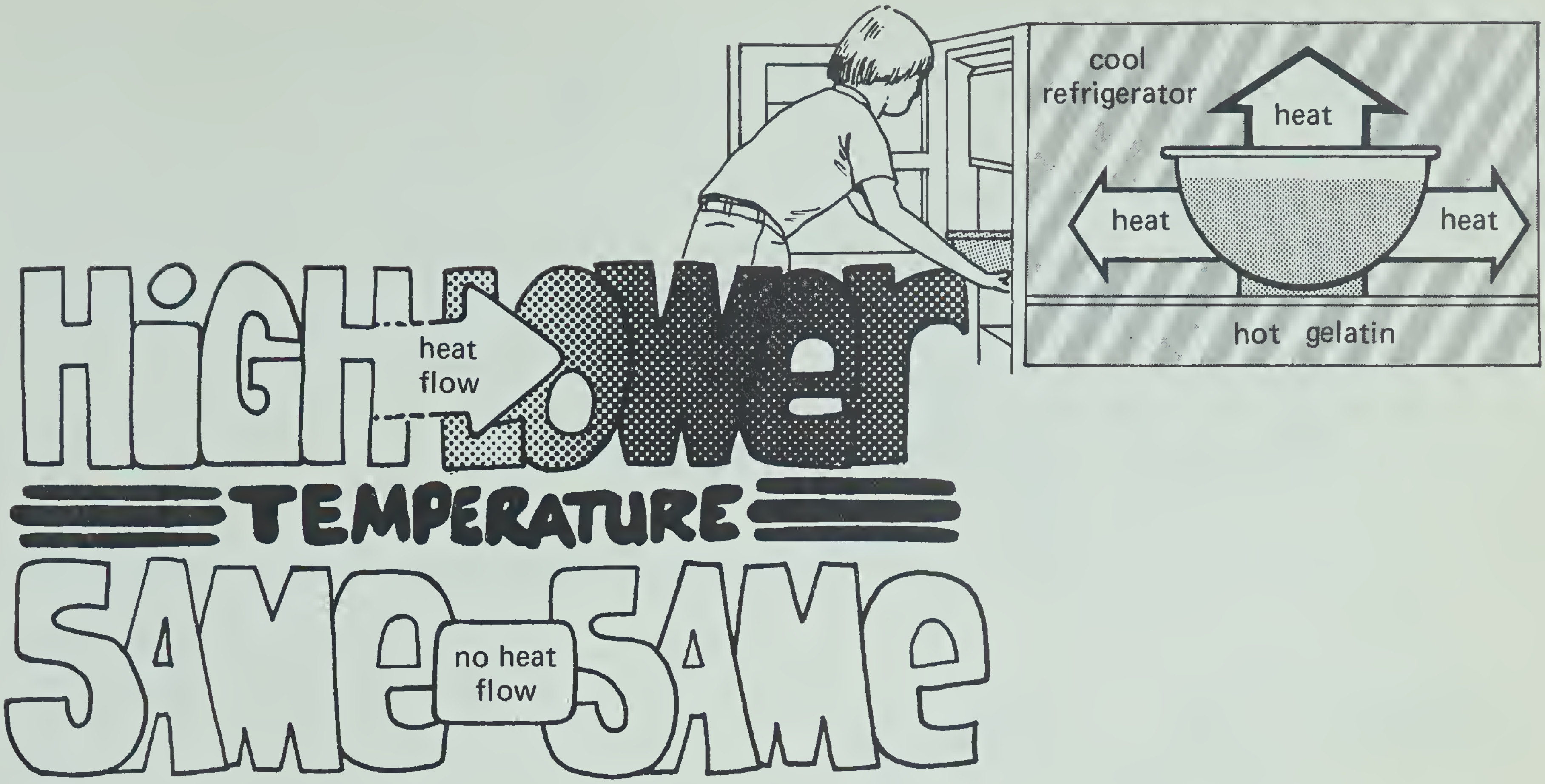


Figure 4-1

● 4-2. Look at Figure 4-1 above. When two substances, such as air and gelatin dessert, are at different temperatures, which way does heat flow?

4-2. From the substance at the higher temperature to the substance at the lower temperature

Heat always flows from a substance at a higher temperature to a substance at a lower temperature. When the substances are at the same temperature, no more heat flows between them.

● 4-3. Tell whether the following statement is true or false. For the temperature of something to decrease, it must be in contact with something that is colder than it is.

4-3. True

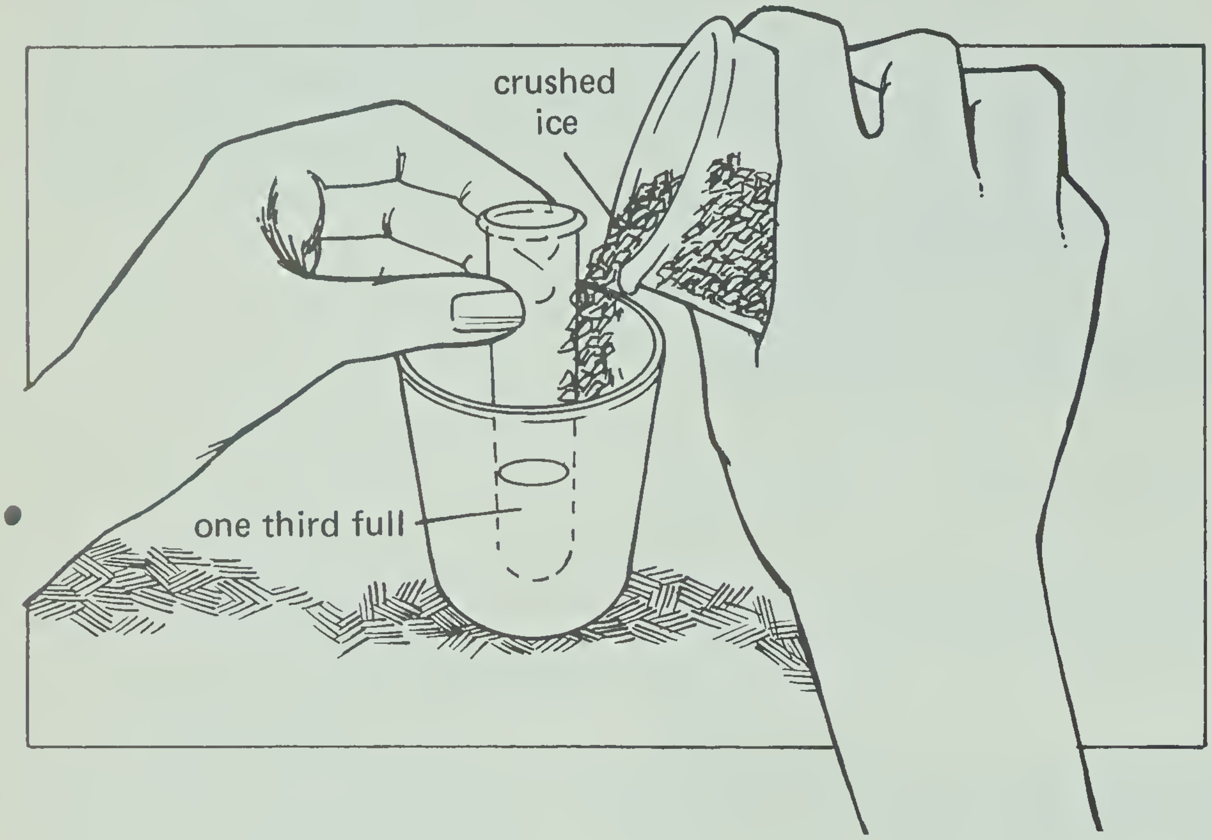
You can apply these ideas by trying to freeze some water. You'll need about thirty minutes and the following materials.

Styrofoam cup
2 thermometers matched for this activity
medium test tube
250-ml beaker filled with crushed ice
watch or clock
safety goggles

If you're not sure how to measure temperature, read "Resource Unit 14: Using a Thermometer" now. Then begin with Step A.

CONTAINER	TEMPERATURE (in °C) FOR TIME PERIOD							
	0	2	4	6	8	10	12	14
Test tube								
Cup								

A. In your notebook, make a table like the one shown here.



B. Fill the test tube about one-third full of tap water, and put it into the cup.

C. Pour crushed ice into the cup around the test tube. Don't get any ice in the test tube.

D. Wait a minute or two for some of the ice in the cup to melt. (This will help you get good contact with the thermometer bulb for an accurate reading.) Measure the temperature of the tap water in the test tube with one thermometer. Then measure the temperature of the melted ice with the other thermometer.

E. Record the temperatures from Step D in your table for Time 0. Then wait two minutes and measure and record the temperatures again for Time 2. Repeat this process until your table is complete.

- 4-4. Which way did heat flow — from the ice to the tap water or the opposite?

4-4. Heat flowed from the tap water to the ice.

- 4-5. What eventually happened to the two temperatures?

4-5. They became the same.

- 4-6. Did the tap water in the test tube ever freeze?

4-6. No

Figure 4-2 below shows what happened during your investigation.

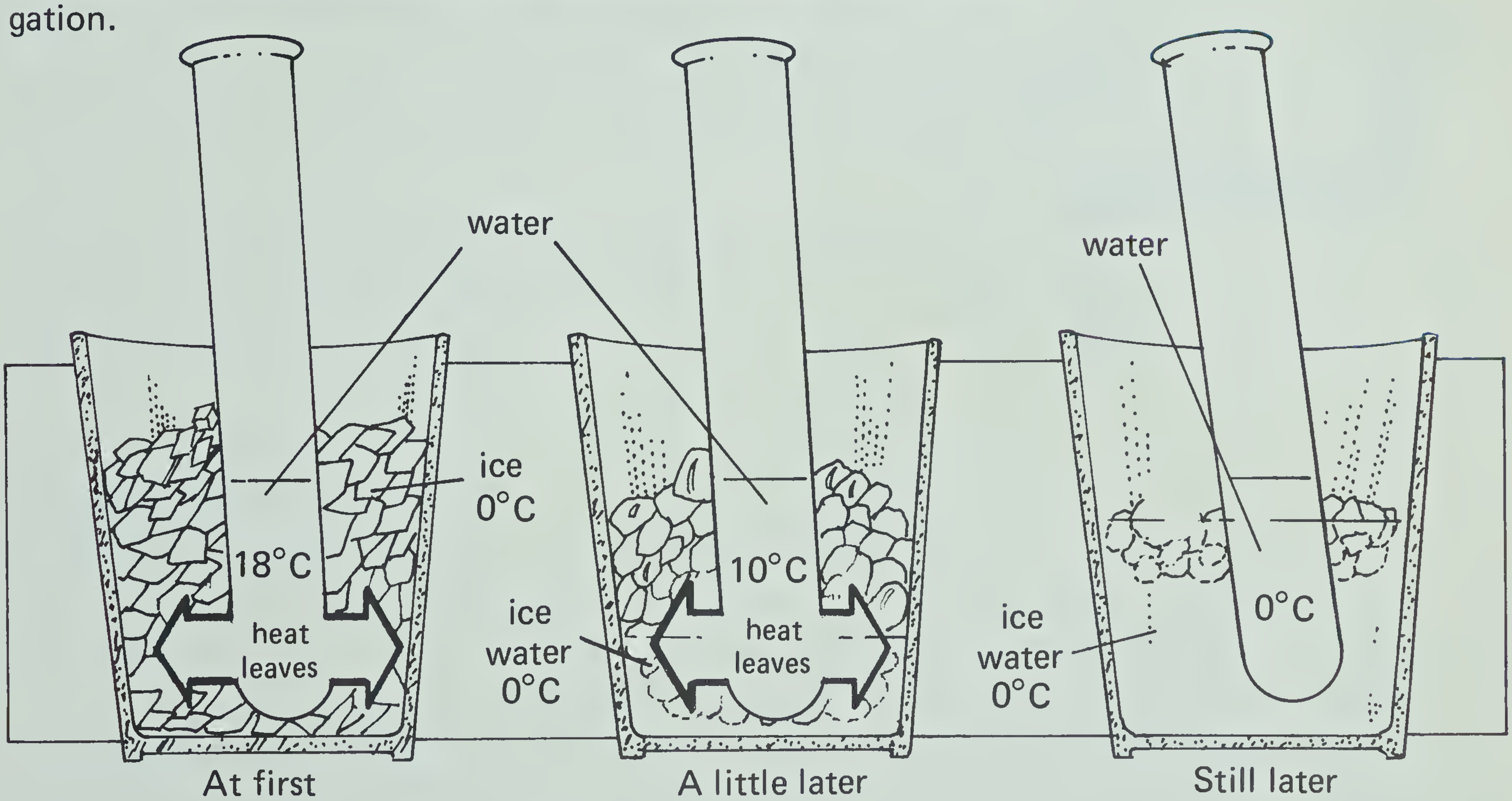


Figure 4-2

- 4-7. Why did the temperature of the tap water get lower?

4-7. Heat left the tap water.

- 4-8. Why did the ice melt?

4-8. Heat entered the ice.

- 4-9. Why didn't the tap water freeze?

4-9. Once the temperatures were the same, no more heat flowed. (Water at 0°C won't freeze unless more heat is taken away.)

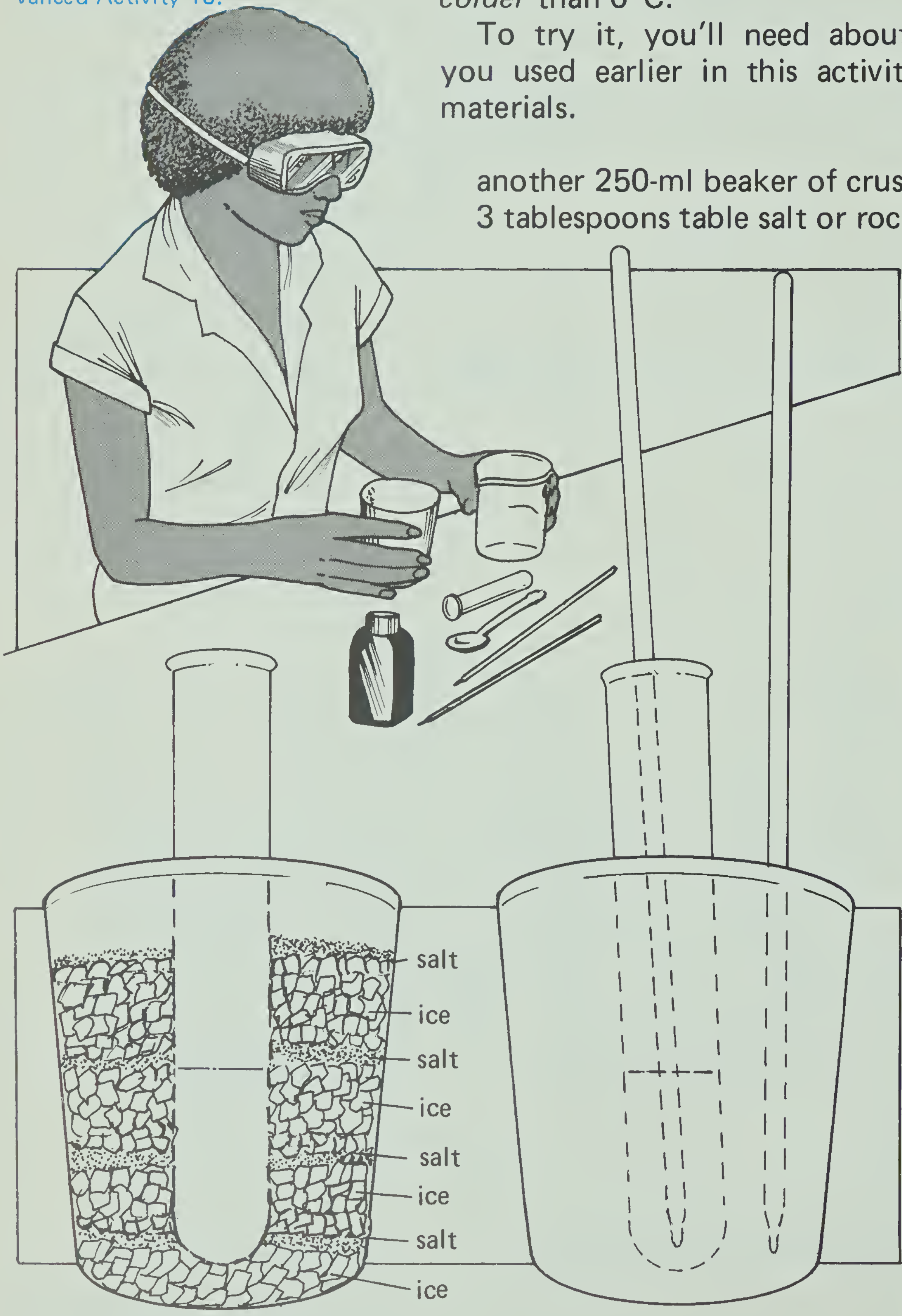
Salt lowers the freezing point for water. This principle, which is also used to clear ice from roads and sidewalks, is covered in advanced Activity 13.

You might have thought that the temperature of the ice water would increase when heat entered it. The reason it didn't is that the heat was used up in melting the ice.

You didn't get the tap water in the test tube to freeze by just putting it into ice. That just lowered its temperature to 0°C. To get water at 0°C to freeze, you must take still more heat from it. You'll have to surround the test tube with something that is *colder* than 0°C.

To try it, you'll need about fifteen minutes, the materials you used earlier in this activity, and the following additional materials.

another 250-ml beaker of crushed ice
3 tablespoons table salt or rock salt



A. Fill the test tube about one-third full of water. Put the test tube into the cup.

B. Add a thin layer of crushed ice to the cup around the test tube. Sprinkle a light layer of salt on top of the layer of ice.

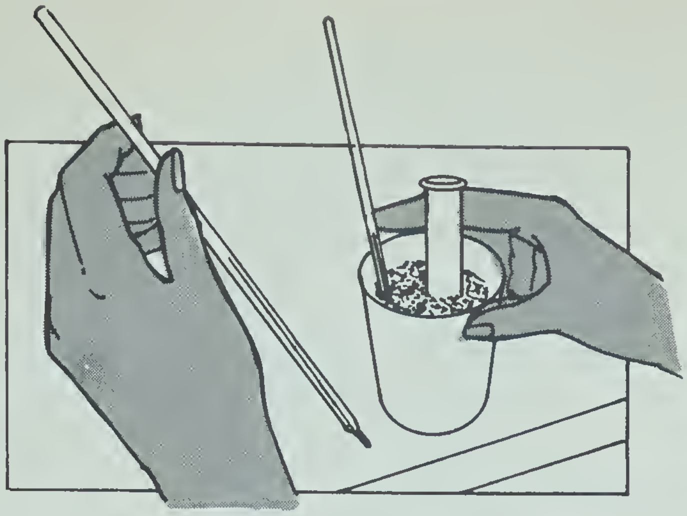
C. Continue to add layers of ice and salt until the cup is full.

D. Measure the temperature of the ice–salt mixture with one thermometer, making sure that the ice has good contact with the thermometer bulb. Measure the temperature of the water in the test tube with the other thermometer.

4-10. [Answers will vary but should be below 0°C.]

- 4-10. What is the temperature of the ice–salt mixture?

E. Remove the thermometer from the test tube. Watch the temperature of the ice–salt mixture for a few minutes. Also observe what happens in the test tube.



Take thermometer out of test tube.

Be sure students remove the thermometer, since it may break if it is frozen into the water.

• 4-11. Is the temperature of the ice–salt mixture above, below, or at 0°C when the tap water in the test tube freezes?

4-11. Below 0°C [between –5°C and –10°C]

☆ **4-12. To freeze ice cream, must the temperature of the ice–salt mixture be above, below, or the same as the freezing temperature of the ice cream? Explain your answer.**

4-12. It must be below the freezing temperature of the ice cream so that heat will be able to flow from the ice cream as it freezes.

You may want to further clarify the idea that liquids won't freeze until the heat of fusion is removed.

Ice cream freezes at about –5°C. So to freeze ice cream, the ice–salt mixture around the ice cream must be at a temperature lower than –5°C.

Now, here's that recipe for ice cream you were promised at the beginning of this activity. The recipe makes about four litres of ice cream. It serves about sixteen adults, so it should be just enough for two or three high-school science students.

IOWA HOMEMADE ICE CREAM

6 eggs
500 cm^3 (2 cups) sugar
25 ml vanilla extract
5 ml almond extract
3 ml lemon extract
2 large cans evaporated milk
about 2 litres whole milk

pinch of table salt
ice cubes
rock salt
ice-cream freezer
large bowl
egg beater or mixer

In a large bowl, beat six eggs until they are fluffy. Then fold in the sugar, vanilla, almond, lemon, salt, and evaporated milk. Mix everything until smooth and creamy. Pour the mixture into the freezer container. Add whole milk to about 5 cm from the top of the container.

Assemble the ice-cream freezer. Use alternate layers of ice and salt around the freezer container. Your ice cream should be ready to eat after about twenty minutes of cranking or mixing.

☆ **4-13. Why won't ice cream freeze in ice alone? Why is salt added to the ice?**

4-13. The temperature of the ice is not low enough to take away enough heat from the ice cream to freeze it. Salt is added to lower the temperature of the ice below that temperature at which ice cream freezes.

ACTIVITY EMPHASIS: The solubility of solids in liquids usually increases with increasing temperature, whereas the solubility of gases usually decreases.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

Contaminated potassium nitrate can be a fire and explosion risk if the solid is subjected to heat or shock. Have students weigh it in a beaker or watch glass rather than on paper.

ACTIVITY 5: HOT AND COLD SOLUTIONS

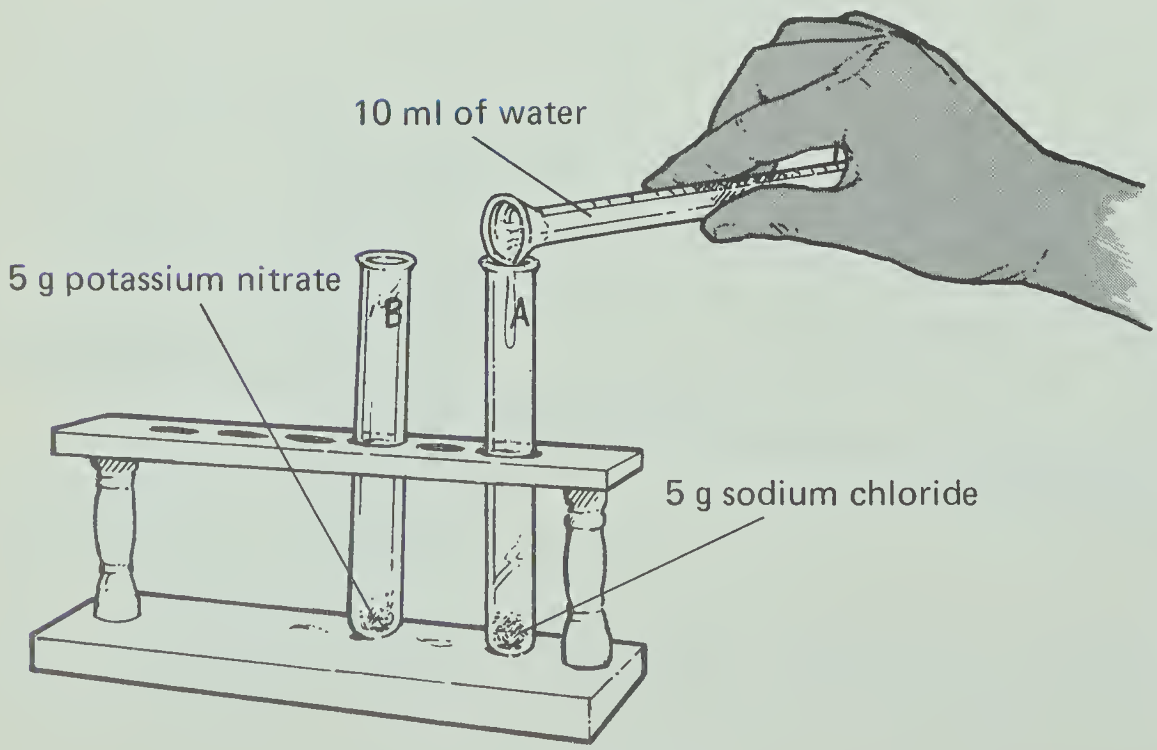
You know that sugar dissolves more easily in hot liquids than in cold ones. It's also true that more sugar will dissolve in a hot liquid than in the same amount of the liquid when it's cold.

You can find out whether this is true for substances other than sugar. You'll need about twenty minutes and the following materials.

balance
10-ml graduated cylinder
2 medium test tubes with stoppers
test-tube rack
5 g potassium nitrate, KNO_3
5 g sodium chloride, NaCl
250-ml beaker
Bunsen burner or other heat source
grease pencil
paper towel
ring stand, ring, and wire gauze
safety goggles

If you're not sure how to use a balance, read "Resource Unit 10: Using Balances." And if you need some help in measuring volumes, read "Resource Unit 5: Measuring Volume." You might also check "Resource Unit 17: Using a Burner." Then return to Step A.

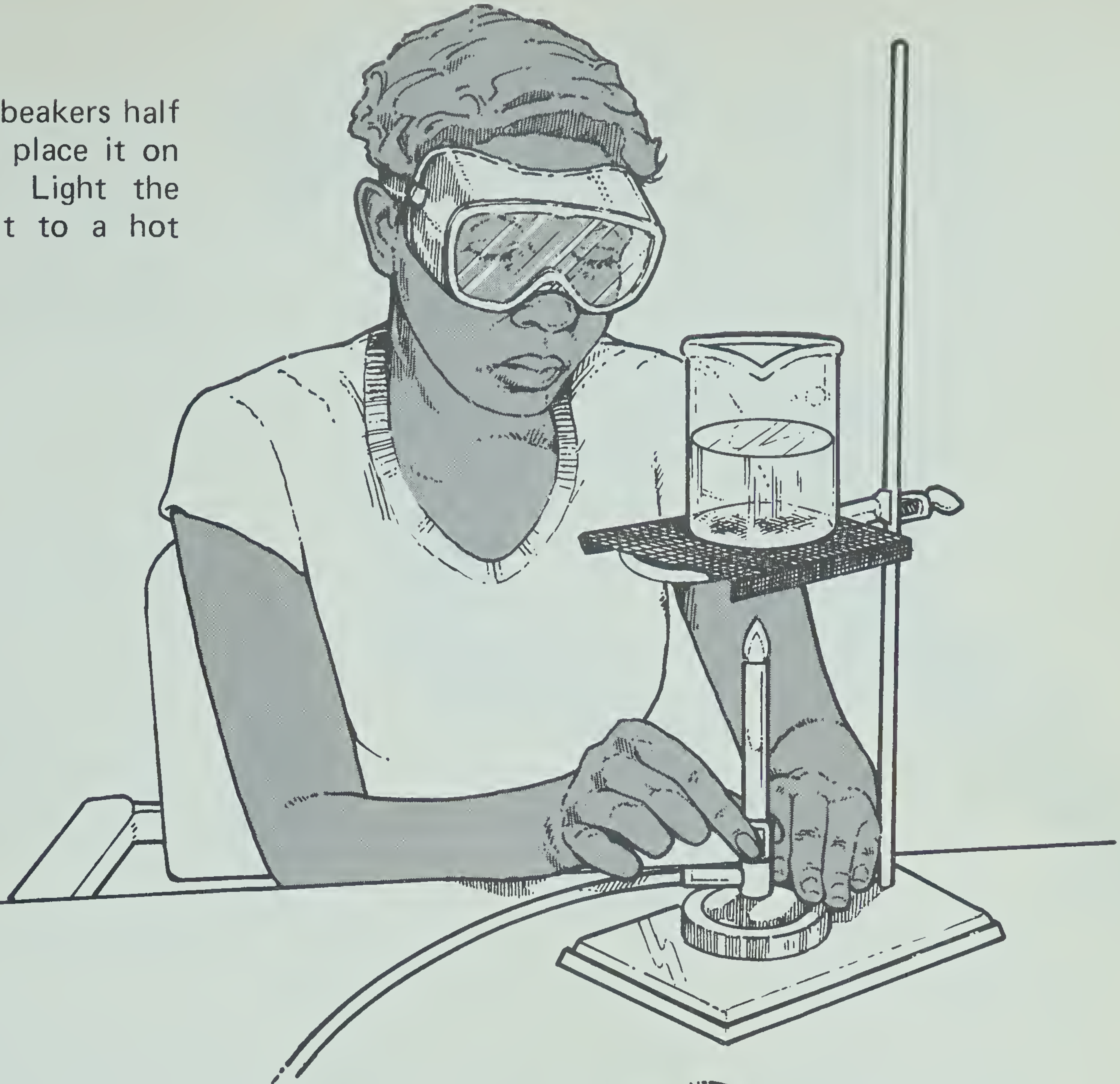
CAUTION	Remember to wear your safety goggles.



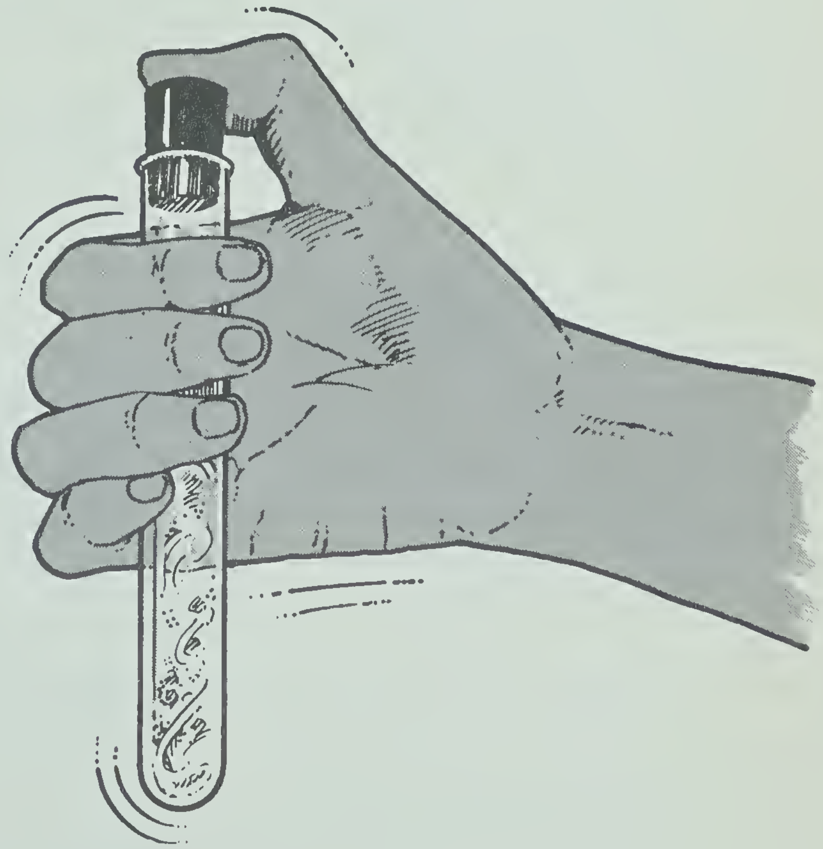
A. With the grease pencil, label the test tubes *A* and *B*. Put both test tubes in the rack. Add 5 g of sodium chloride, NaCl, to Test Tube A. Add 5 g of potassium nitrate, KNO_3, to Test Tube B.

B. Use the graduated cylinder to add 10 ml of water to each test tube. With the grease pencil, mark the level of the solid in each test tube.

C. Fill one of the beakers half full of water and place it on the ring stand. Light the burner and set it to a hot flame.



D. Stopper both test tubes. Shake each one for about three minutes to dissolve as much solid as possible. Mark the levels of the solids again.



• 5-1. Did all of the sodium chloride dissolve? All of the potassium nitrate?

5-1. No; no

E. When the water in the beaker begins to boil, turn off the burner. Remove the stoppers, and put the two test tubes into the beaker. Leave them in the water for about one minute.

F. Stopper the test tubes. With paper towel around the top, gently turn each test tube upside down several times. Put the test tubes back into the rack. Note the levels of the solids now.

Long, needlelike crystals of KNO_3 can be grown by cooling the hot solution and then adding a seed crystal.

5-2. Yes; all of the potassium nitrate dissolved. (The difference in solubility of NaCl at 20°C and 100°C is too small to be reliably noted by most students.)

5-3. Potassium nitrate

5-4. At 75°C; more of a solid usually dissolves at higher temperatures.

● 5-2. Did more of either solid dissolve when the solutions were warmed? Which one?

● 5-3. Which has greater *solubility* (dissolves more) in warm water — sodium chloride or potassium nitrate?

Scientists have done many experiments like this with many different solids and liquids. They have found that, in general, the higher a liquid's temperature, the more of a particular solid it will dissolve.

☆ **5-4. In general, will more of a solid dissolve in 100 ml of water at 25°C or at 75°C? Why?**

When a gas dissolves in a liquid, it's a different story. To see this, you'll need about thirty minutes and these materials.

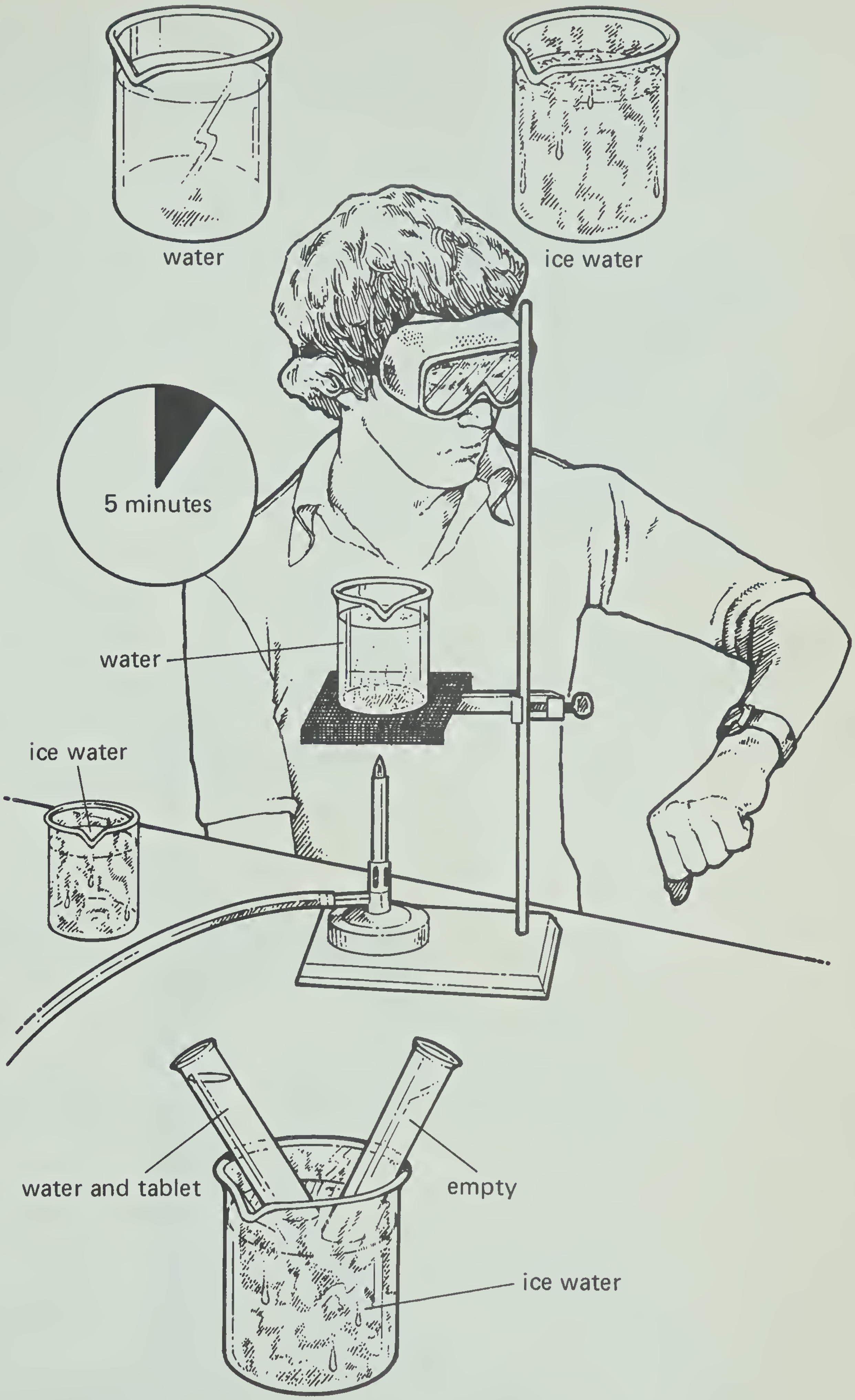
¼ of an Alka-Seltzer tablet
2 250-ml beakers
2 medium test tubes
crushed ice
Bunsen burner or other heat source
ring stand, ring, and gauze
safety goggles

A. Fill one beaker almost full of crushed ice and add some water to it. Fill the other beaker almost full of water.

B. Put the beaker with the water on the stand. Light the burner, and heat the water.

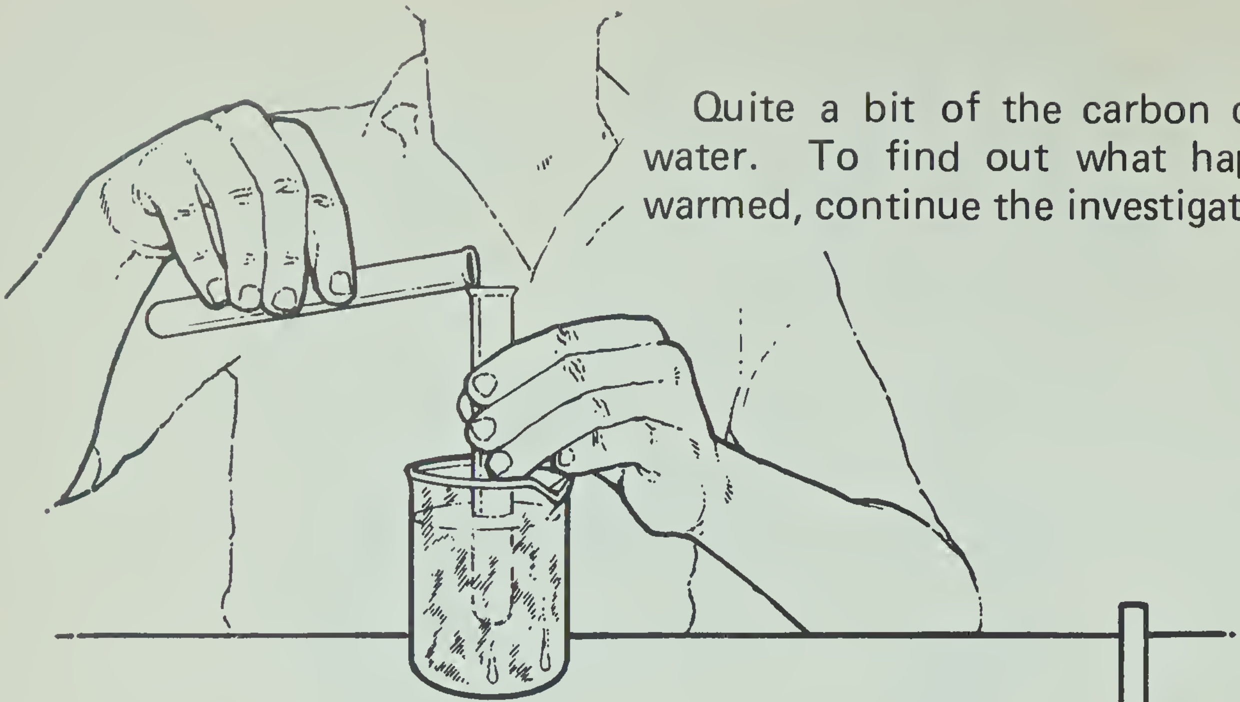
C. Wait five minutes for the water in one beaker to get hot and the water in the other beaker to get cold. Turn off the burner.

D. Fill one test tube almost full with cold water from the ice-filled beaker. Drop in ¼ Alka-Seltzer tablet. Put the test tube into the ice-filled beaker to keep it cold. Put the other test tube into the ice-filled beaker to get cold too.



● 5-5. When you dropped the tablet into the cold water, it fizzed and produced carbon dioxide gas. Some of the gas escaped. Where did the rest of it go?

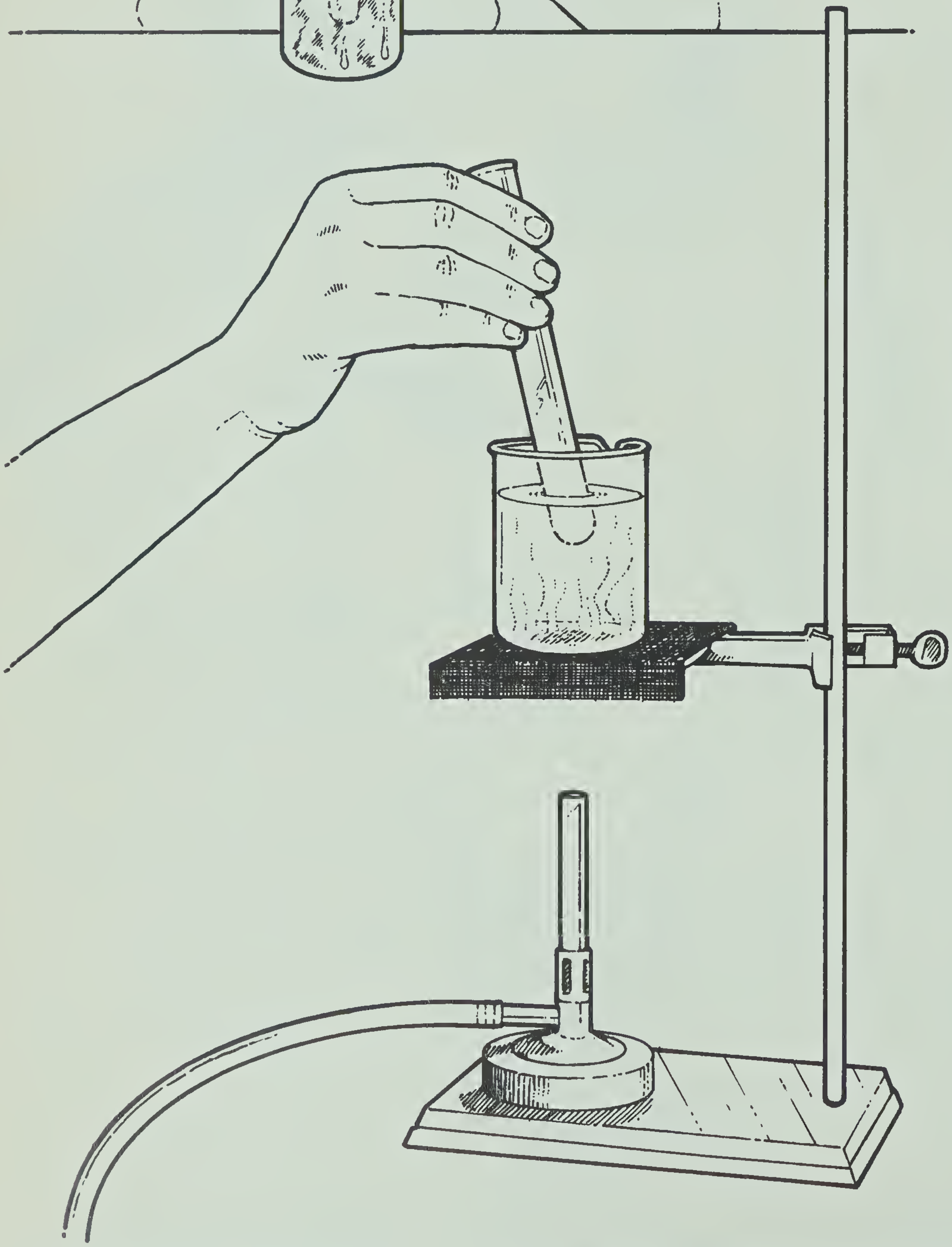
5-5. It dissolved in the water.



Quite a bit of the carbon dioxide gas dissolved in the cold water. To find out what happens when the cold solution is warmed, continue the investigation.

E. After the tablet is almost finished fizzing, pour most of the clear liquid from the test tube into the other cold test tube. Set the first test tube aside. You won't need it any more.

F. Put the bottom of the test tube into the hot water for a few seconds and watch what happens. Then put it back into the ice-filled beaker to keep it cold.



5-6. Small gas bubbles (carbon dioxide) formed in the solution.

● 5-6. What happened when you put the bottom of the test tube into the hot water?

The bubbles you observed in Step F (page 24) were carbon dioxide gas. It isn't as soluble in hot water as in cold. When the cold solution was warmed, the gas came out of solution, forming bubbles.

5-7. The gas forms bubbles and leaves. Less carbon dioxide stays in solution.

5-8. 25°C; usually, more gas dissolves at a lower temperature.

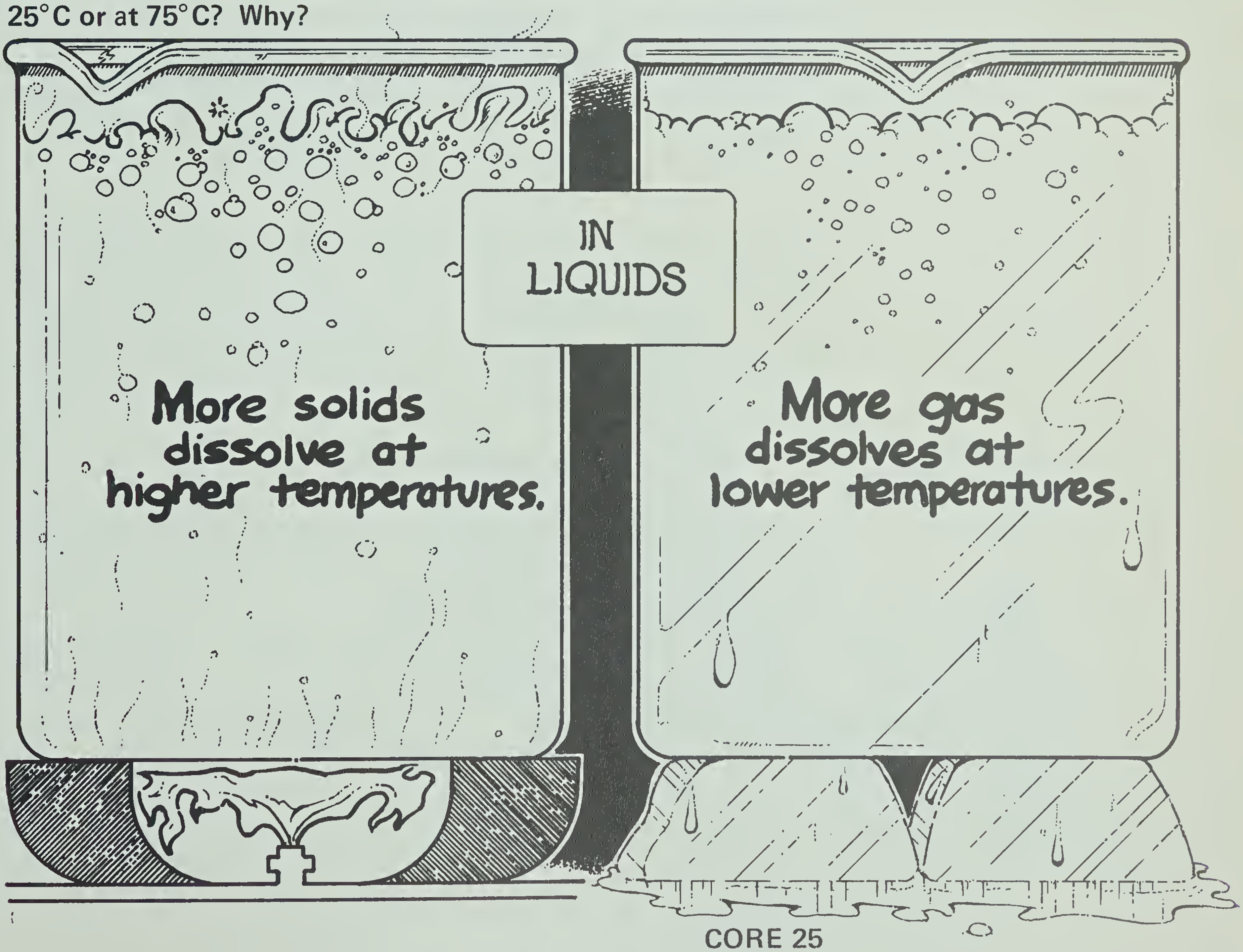
G. Repeat Step F.

"Carbonated" soft drinks contain dissolved carbon dioxide gas. The gas helps give these drinks their biting taste.

● 5-7. What happens to the dissolved carbon dioxide gas in cold soda pop as it is warmed?

Investigations of other gases and liquids show similar results. Usually, the higher the temperature of a liquid, the less gas dissolves in it.

☆ **5-8. Would more of a gas usually dissolve in 100 ml of water at 25°C or at 75°C? Why?**



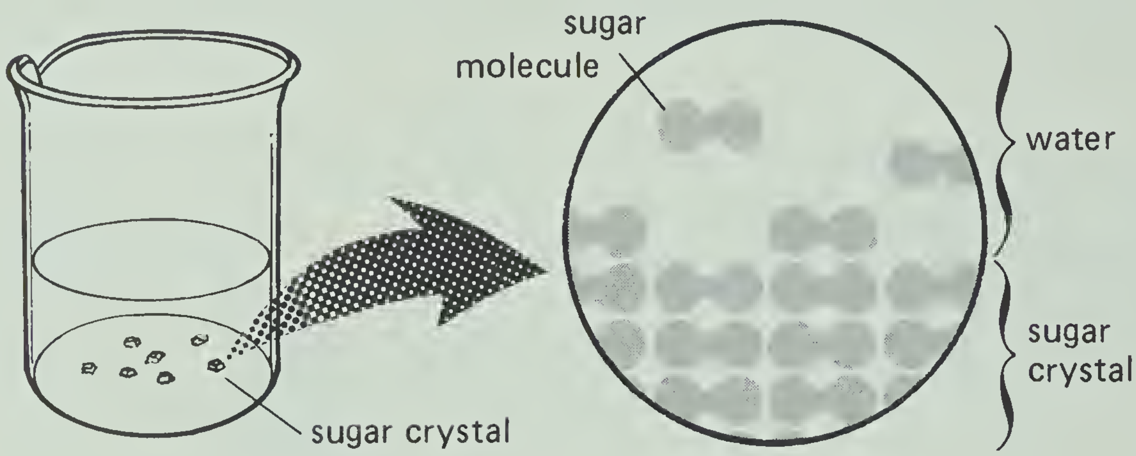
ACTIVITY EMPHASIS: Saturated and supersaturated sugar solutions are made by heating and cooling sugar and water. Stirring, cooling, and evaporation affect the size of sugar crystals in candy making.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

ACTIVITY 6: CANDY AND CRYSTALS

Most candy is made by dissolving sugar in water and then cooking the mixture. Often flavorings, nuts, and fruits are added. In this activity you'll make some unflavored candy called *fondant* [FON-dunt]. Some people also call it *white fudge*. You'll find a recipe for another unflavored candy — rock candy — at the end of this activity. It's easy to make too, but it takes more than a week to be ready.

When you dissolve sugar in water, molecules of sugar get loosened from the sugar crystals and mixed up with the water. (If you're not sure what a molecule is, read "Resource Unit 24: Atoms, Molecules, and Ions.")



As long as more and more sugar continues to dissolve, the solution is said to be *unsaturated*. When no more sugar will dissolve at that temperature, the solution is called *saturated*. This applies to all substances other than sugar too. A substance's *solubility* in water is the amount it takes to make a saturated solution at a certain temperature.

6-1. A saturated sugar solution is one in which no more sugar will dissolve at that temperature.

☆ **6-1. What is a saturated sugar solution?**

The hotter the water is, the more sugar can be dissolved. Look at Figure 6-1 below. It shows the solubility of sugar in water at two different temperatures.

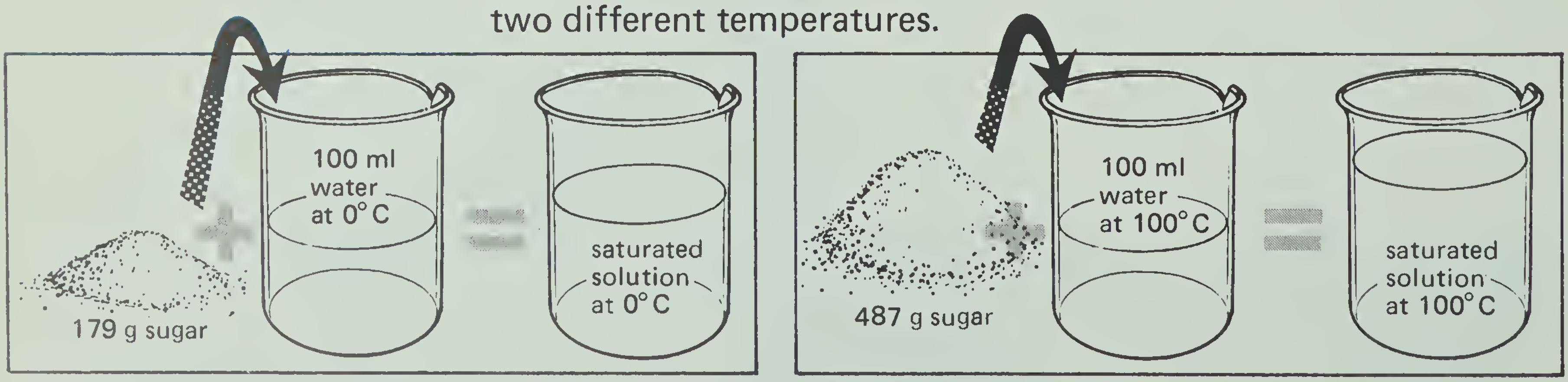


Figure 6-1

• 6-2. Which contains more dissolved sugar — a cold saturated solution or a hot saturated solution? (See Figure 6-1, page 26.)

6-2. A hot saturated solution contains more sugar. The solubility of sucrose (cane sugar) in water at 0° C is 179 g/100 ml. At 100° C, it is 487 g/100 ml.

Be sure students thoroughly clean all equipment and wash their hands before making the fondant.

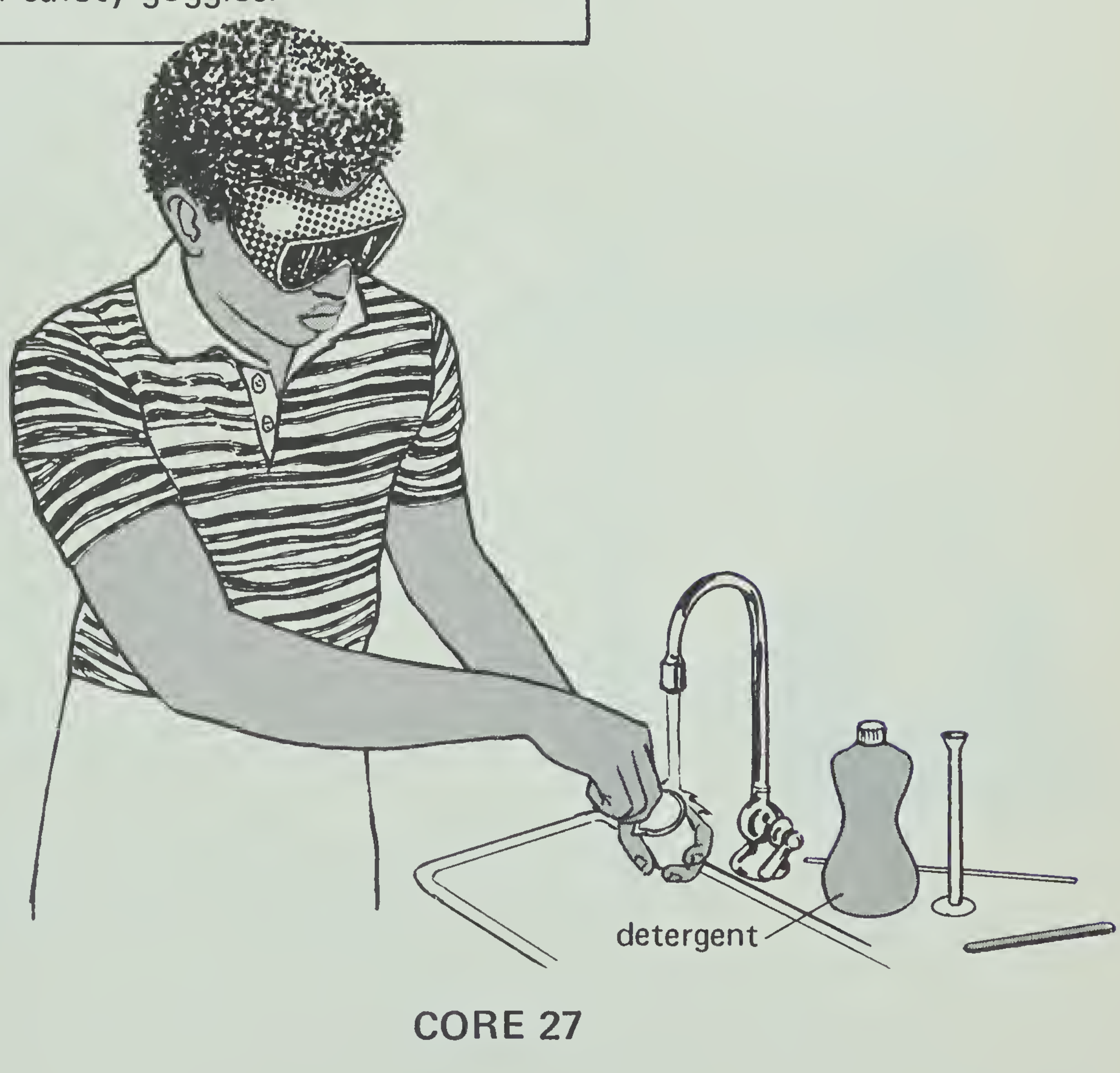
You can see what happens when you cool a hot saturated sugar solution and make some fondant at the same time. You'll need about forty-five minutes and these materials.

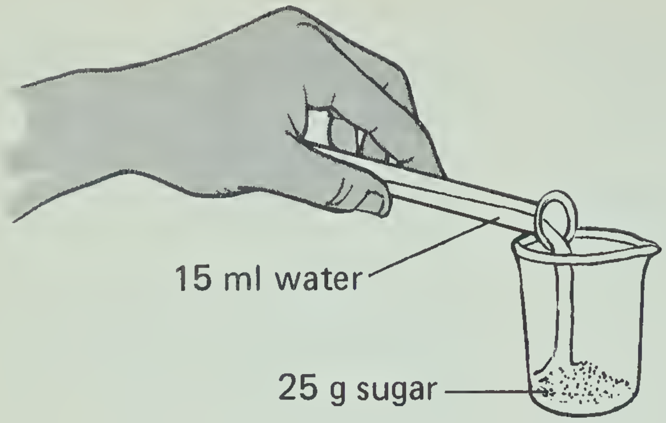
100-ml beaker
thermometer with scale markings to at least 114°C
Bunsen burner or other heat source
ring stand, ring, and gauze
wooden tongue depressor
paper towel
10-ml graduated cylinder
clamp and slotted cork to hold thermometer
25 g sugar
safety goggles

Before you begin Step A, you might want to check "Resource Unit 5: Measuring Volume" and "Resource Unit 10: Using Balances." "Resource Unit 14: Using a Thermometer" and "Resource Unit 17: Using a Burner" will also be helpful.

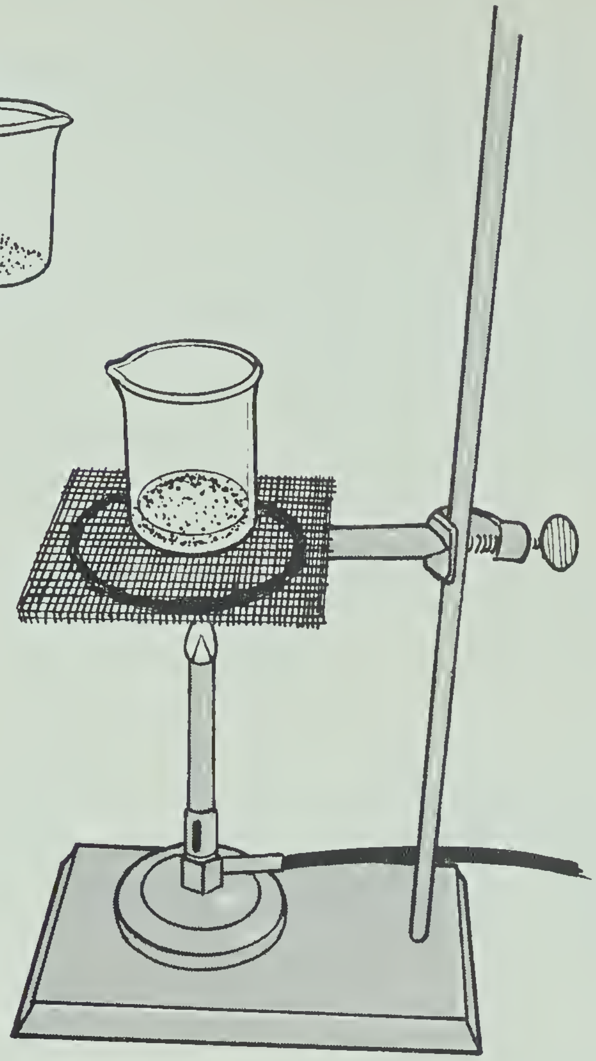
CAUTION	Wear your safety goggles.

A. Wash the beaker, thermometer, graduated cylinder, and tongue depressor with hot water and detergent, scrubbing them thoroughly. Be sure they're very clean, since you'll be eating the candy you make. Wash your hands well also.

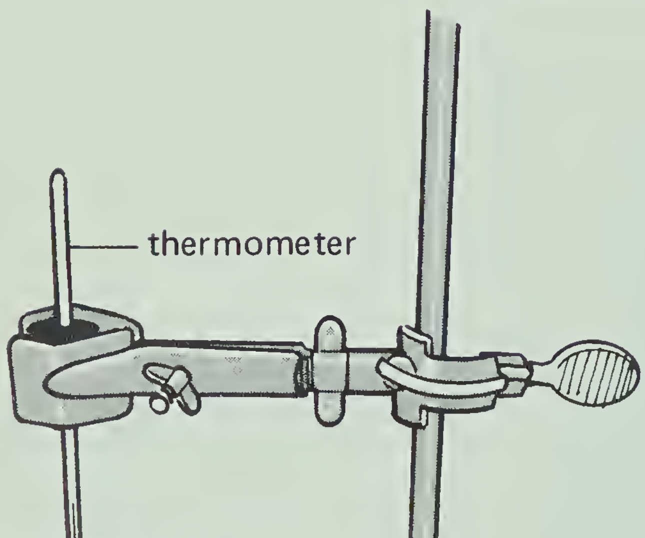




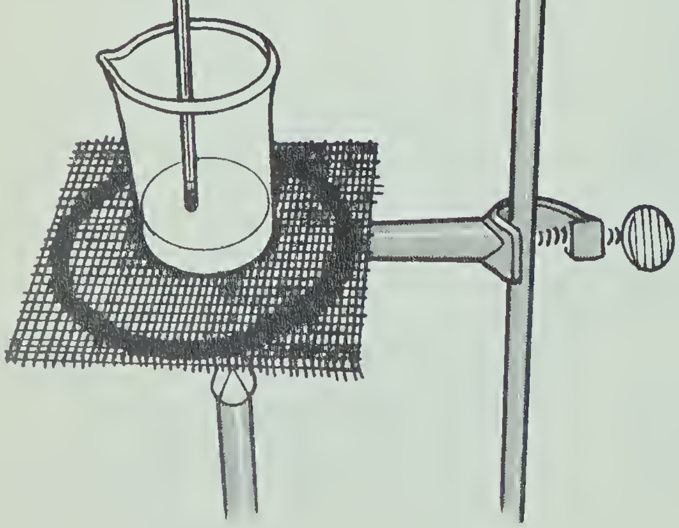
B. Put 25 g of sugar into the beaker. Add 15 ml of water.



C. Light the burner, and gently heat the contents of the beaker to boiling. With the tongue depressor, scrape any sugar crystals from the sides of the beaker into the hot liquid. Make sure that all the sugar dissolves. If you have to stir the liquid, do so gently. Don't splash it.

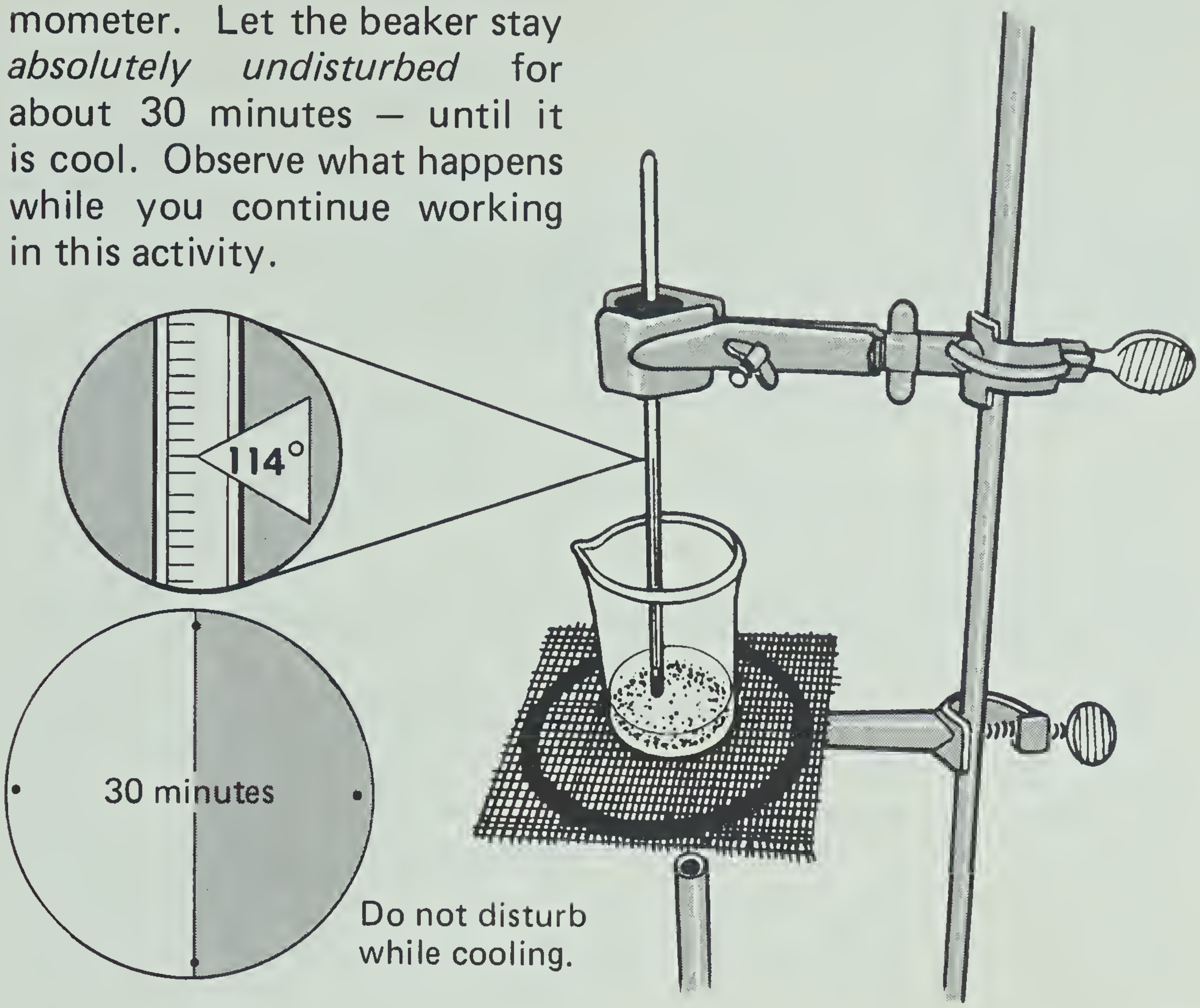


D. Put the thermometer into the slotted cork. Clamp them into position so that the bulb is just above the bottom of the beaker.



E. Keep heating the mixture gently. Use the thermometer to measure the temperature of the boiling liquid.

F. When the temperature reaches 114°C, turn off the burner, and remove the thermometer. Let the beaker stay *absolutely undisturbed* for about 30 minutes — until it is cool. Observe what happens while you continue working in this activity.



When the solution reached 114°C and you turned off the burner, it was a saturated solution. As the solution cools, it becomes what is called *supersaturated*. More sugar remains dissolved in it than would be expected to at the lower temperature.

The proper temperature will be lower than 114°C at higher elevations. See "Advance Preparations" in ATE front matter (page ATE 6) for details.

Not all substances exhibit supersaturation in water. Excess solid usually crystallizes out while cooling.

• 6-3. Describe the appearance of the supersaturated sugar solution as it cools.

6-3. It's a clear liquid.

You may have thought that the extra sugar would recrystallize (come out of solution) as the solution cooled. But if you've been careful, and the beaker is clean and smooth, the extra sugar stayed in solution and you got supersaturation.

☆ **6-4. How would you make a supersaturated sugar solution?**

6-4. By carefully cooling a hot saturated solution

A supersaturated solution won't stay that way — it's very unstable. The extra dissolved substance will always come out of solution eventually. But sometimes the process takes a long time.

That's how you can make rock candy. If you let the supersaturated solution cool with a string in it, the extra sugar would eventually form solid sugar crystals on the string. Figure 6-2 below shows the process.

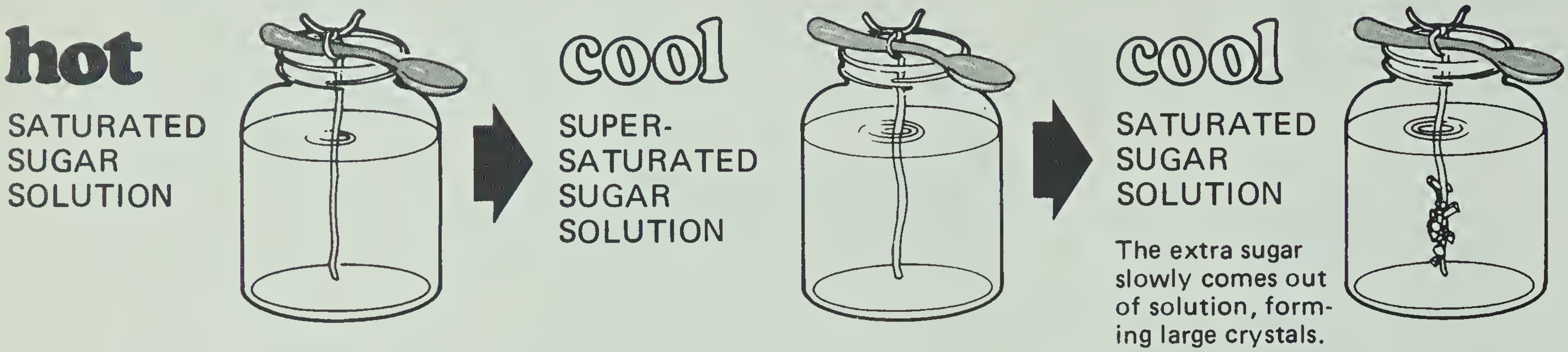


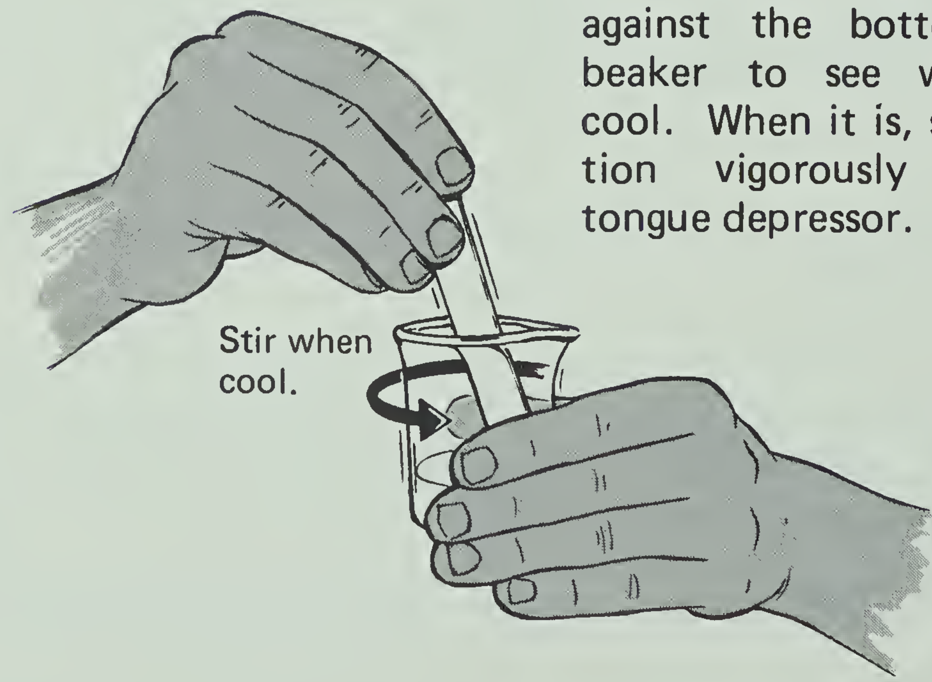
Figure 6-2

6-5. A hot saturated solution cools to make a supersaturated solution, from which large crystals eventually form.

☆ **6-5. How are the large sugar crystals in rock candy formed?**

When all the extra sugar that made the solution supersaturated recrystallized, you'd have two or three dozen large crystals. The rest would remain dissolved in the now simply saturated solution. If you wanted to get still more large crystals, you'd have to let the solution stand longer so that more water would evaporate. But you don't have to wait for rock candy crystals to form in your supersaturated solution. You can make fondant much more quickly.

G. Put your hand very gently against the bottom of the beaker to see whether it's cool. When it is, stir the solution vigorously with the tongue depressor.



The investigation will work even if the solution is somewhat warm.

6-6. The solution turned creamy white and then solidified.

● 6-6. Describe what happened as you stirred the solution.

When you stirred the cool supersaturated solution, all the extra sugar molecules came together almost immediately to form crystals.

● 6-7. Taste a little of your fondant candy. What is its texture?

6-7. Creamy and smooth

The fondant candy you made by this method has thousands and thousands of very tiny sugar crystals. Since the sugar crystals are so very small, the candy has a creamy, smooth texture.

☆ **6-8. How are the sugar crystals in smooth fondant formed?**

6-8. A hot saturated sugar solution is cooled to form a cool supersaturated sugar solution, which is then stirred vigorously to form tiny crystals.

If you had stirred the sugar solution before it was cool enough, the fondant wouldn't be smooth. The extra sugar molecules would have come together a few at a time, forming larger crystals. So the fondant would be sugary and would feel gritty. (If you've got the time and the sugar, you might want to repeat Steps A through E and then stir the solution vigorously as it cools.)

● 6-9. In making fondant, why should you cool the supersaturated sugar solution before you stir it?

6-9. So that you get tiny crystals and smooth, not gritty, fondant

Porous cotton string should be used so that evaporation from the string causes the first small crystals to form.

Here's that recipe for rock candy, in case you want to make some at home.

Rock Candy

125 ml (½ cup) water
250 cm^3 (1 cup) sugar

large baby-food jar
piece of cotton string
plastic spoon

Heat the water almost to boiling. While heating it, stir in the sugar until it is all dissolved. Bring the solution to boiling, and boil it for one minute.

Tie a piece of string onto the spoon handle so that one end is long enough to reach nearly to the bottom of the baby-food jar. Wet the string, and suspend the spoon over the top of the jar.

Pour the hot sugar solution into the baby-food jar. Let the whole setup stand for at least a week in a clean location where it won't be disturbed.

Sugar crystals will form on the string as water evaporates from the solution. When they get big enough, you can pull out the string, break off some of the crystals to eat, and put the string back to make more rock candy.

ACTIVITY EMPHASIS: Odors are caused by gases that can be sensed as they evaporate and sublime and diffuse from their es. Baking soda controls by reacting with the acidic that cause them. Activated coal adsorbs odorous gas olecules.

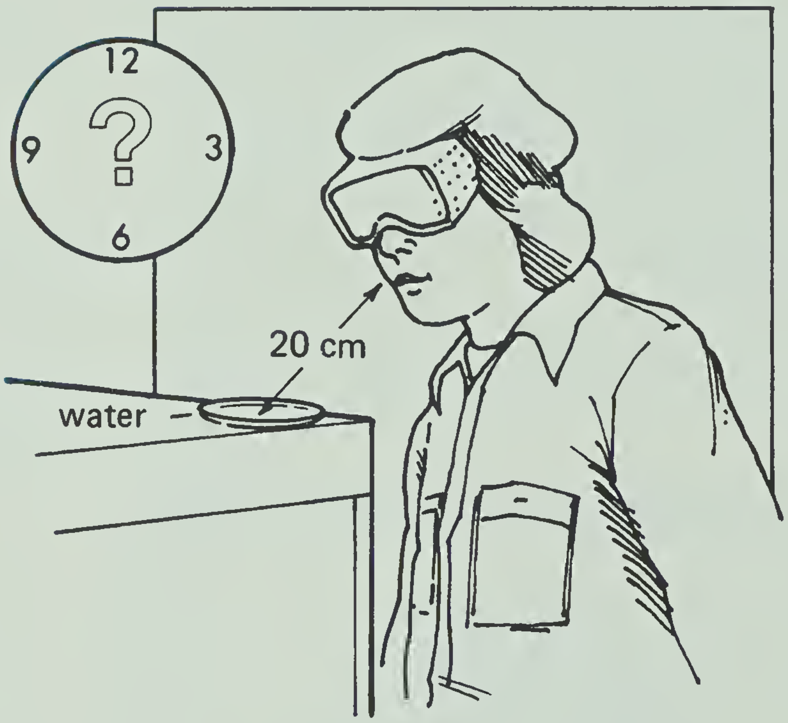
IATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

ACTIVITY 7: GETTING RID OF ODORS

Some odors are pleasant and others aren't, but they certainly are all around. Knowing what causes odors and how they travel can help you get rid of unpleasant ones.

Odors are caused by gases, but not all gases have odors. A simple investigation can illustrate this idea. You'll need the following materials.

2 watch glasses
household ammonia in dropping bottle
water in dropping bottle
clock or watch
ruler or metre stick
safety goggles

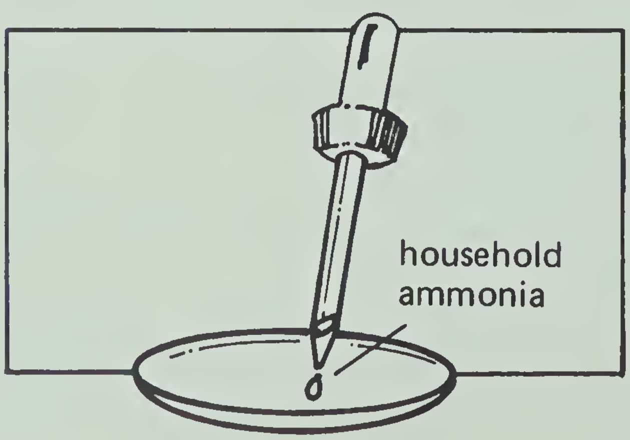


A. Put several drops of water into one of the watch glasses.

B. Position yourself so that your nose is about 20 cm from the watch glass.

C. Time how long it takes for you to smell the water.

At this point, you may be saying to yourself, "Water? I can't smell water!" It's true – you can't.



D. Repeat Steps A, B, and C, using household ammonia instead of water.

7-1. [Answers will vary.]

● 7-1. How long did it take for you to smell the ammonia?

This investigation illustrates one difference between water and ammonia. Ammonia as a gas has an odor, but water doesn't. If you can smell a gas, it has an odor; if you can't, it doesn't.

☆ **7-2. How can you tell whether a particular gas has an odor?**

7-2. If a gas has an odor, it can be smelled.

The ammonia gas you smelled got to your nose from the dish by a process called *diffusion*.

- 7-3. What is diffusion?

7-3. Diffusion is the movement of particles of a substance from a region where the particles are highly crowded to a region where the particles are less crowded.

If you had trouble answering Question 7-3 above, read "Resource Unit 12: Diffusion." Then return to Question 7-4 below.

☆ **7-4. Look at Figure 7-1 below. How does an odor travel from its source to your nose?**

7-4. Molecules of the substance leave the source and diffuse through the air to your nose.

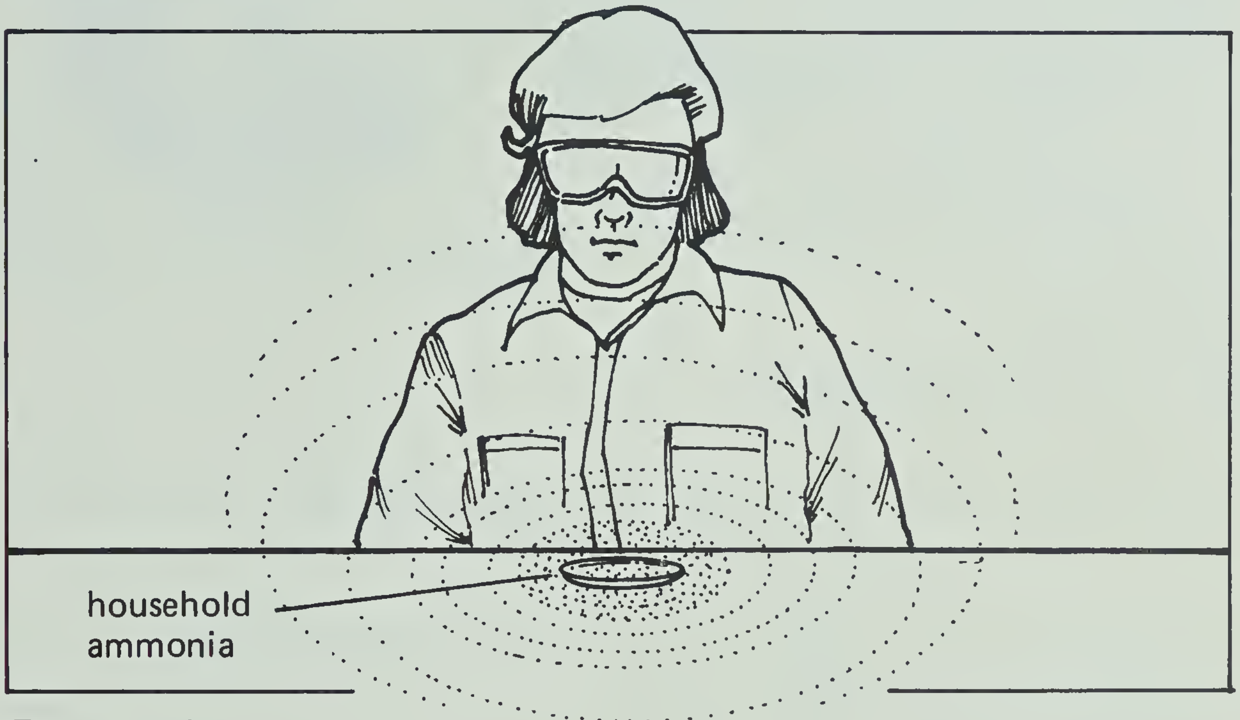
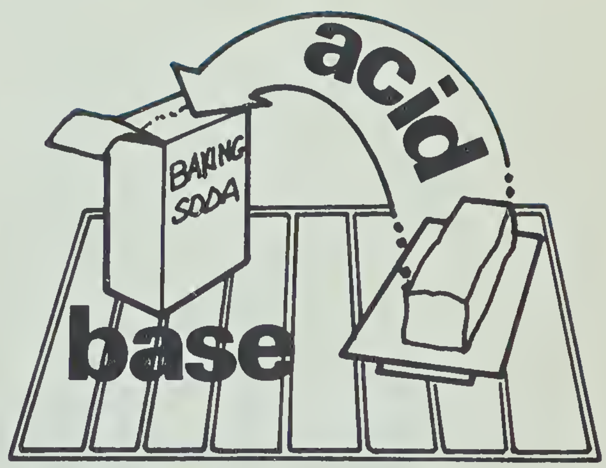


Figure 7-1

Many air fresheners contain substances that mask odors rather than react with them.

The hydrogen carbonate ion in baking soda can also react with bases, such as amines.



Many substances are made up of molecules. If the molecules of a substance are changed by a chemical reaction, its odor is changed too.

Many of the unpleasant odors in a refrigerator, for example, are caused by acids. Acids will react chemically with bases, changing the molecules and thus the odor. You can put a little baking soda (a base) into a refrigerator, and it will react chemically with the acidic gases and improve the smell. You should stir the baking soda occasionally to expose fresh molecules to the acidic gases.

☆ **7-5. Rancid butter odor is caused by butyric acid. How would baking soda help remove that odor from a refrigerator?**

7-5. Baking soda, which is a base, will react chemically with the acidic molecules from the butyric acid.

You may need to point out that it is *ad*sorb, not *ab*sorb.

Another way to get rid of odors is to *adsorb* the molecules of a substance. For this, activated charcoal is often used in such things as gas masks and refrigerators. If you looked at activated charcoal under a microscope, you would see that it is full of tiny holes, like a sponge.

But there is a difference between a sponge soaking up water and activated charcoal getting rid of odors by adsorbing gases. Look at Figure 7-2 below.

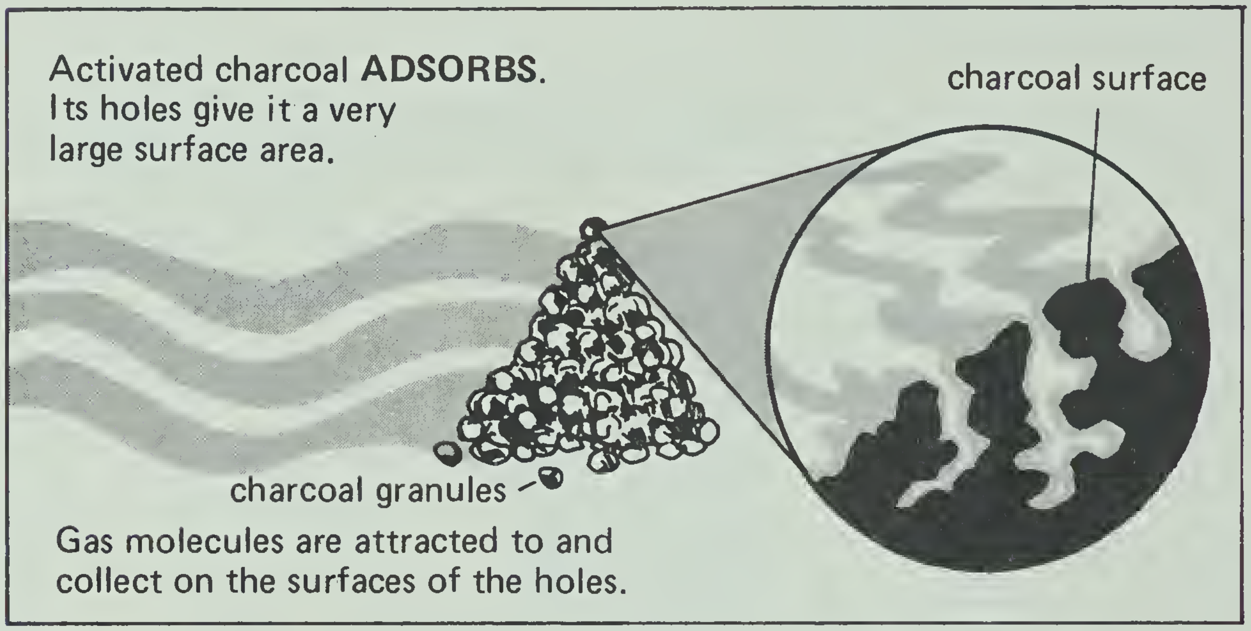


Figure 7-2

7-6. The process in which gas molecules are attracted to and collect on surfaces

Have the "special" test tubes available for this activity.

The student-produced activated charcoal will be used in the last part of this activity.

Have students do this investigation in a well-ventilated area.

☆ 7-6. What is adsorption?

You can make your own activated charcoal. You'll need about twenty minutes and these materials.

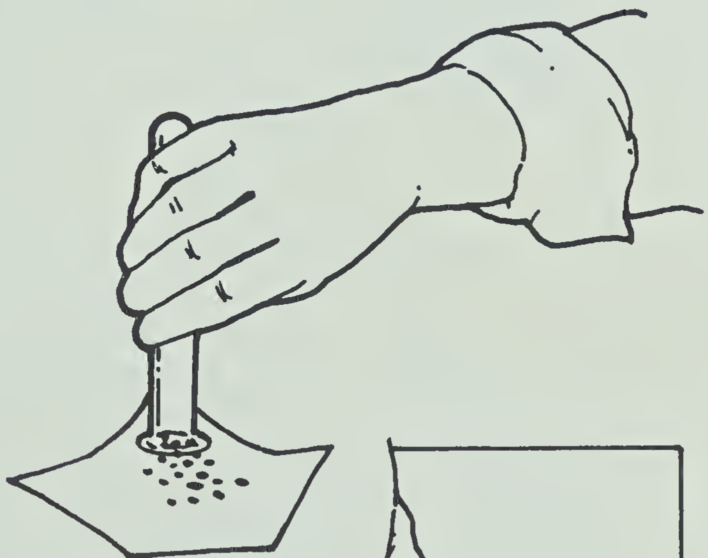
puffed cereal, popped popcorn, or coconut-hull fibers
special test tube for this activity
test-tube holder
Bunsen burner or other heat source
mortar and pestle
safety goggles

CAUTION	Do this investigation in a fume hood or in a well-ventilated area.

A. Fill the special test tube about half full with the puffed cereal, popcorn, or coconut-hull fibers. Using a pencil, pack the material in fairly tightly.

B. Light the burner. Using the test-tube holder, gently heat the bottom of the test tube. After most of the smoke has been driven off, heat the bottom of the test tube more strongly. Stop when the bottom half of the material has been thoroughly charred.

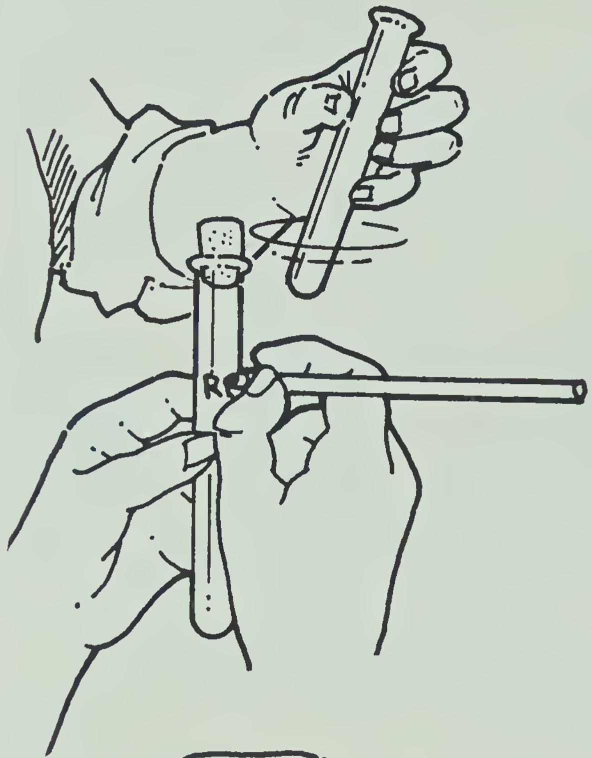
C. Let the test tube cool. Then empty it, and save only the thoroughly charred material. Properly dispose of the rest.



D. Grind the charred material in the mortar with the pestle until most of the pieces are about the size of grains of table salt.

You now have some activated charcoal. To test its effectiveness you'll need a few minutes today and tomorrow and the following materials.

cologne in dropping bottle
2 medium test tubes with stoppers
grease pencil



A. Put one drop of cologne into each test tube. Roll it around the bottom and sides. Then put your activated charcoal into one of the test tubes.

B. Stopper each test tube. Write your initials on them with the grease pencil, and store them overnight.

C. The next day, remove the stoppers and smell the contents of each test tube.

Make provisions for storing these test tubes.

● 7-7. Did the activated charcoal have any effect on the cologne odor? If so, how?

7-7. Yes; the cologne odor is gone from the test tube with the activated charcoal. Gas molecules from the perfume were adsorbed by the activated charcoal.

After your activated charcoal has adsorbed a lot of odorous gas molecules, it becomes deactivated. You can activate it again by reheating it. This drives off the odors it has adsorbed into the air.

☆ **7-8. How does adsorption by activated charcoal help control odors?**

7-8. Odorous gas molecules collect on the surfaces of the activated charcoal.

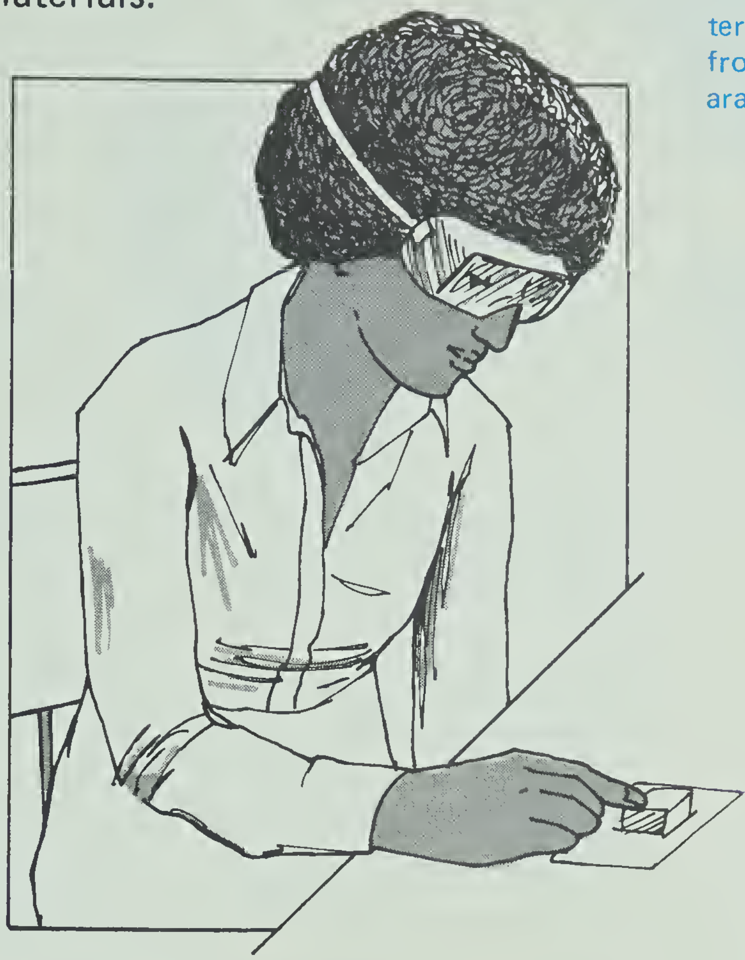
ACTIVITY 8: TENDER MEAT

ACTIVITY EMPHASIS: Meat tenderizers contain an enzyme called *papain* that breaks up long strands of meat protein.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

If you've ever eaten a lot of fresh, raw pineapple, you probably remember a certain feeling inside your mouth. To see what raw pineapple has to do with meat tenderizers, you'll need about ten minutes and the following materials.

cube of gelled gelatin
small piece of aluminum foil
meat tenderizer
small piece of paper
safety goggles



A. Put the cube of gelatin on the piece of aluminum foil. Poke at it a little. Then, think about how you would describe it to someone who has never seen gelatin.

● 8-1. How would you describe gelatin to someone who had never seen any (without using your hands)?

8-1. [Answers will vary and should be interesting.]

A chemist would describe gelatin something like this. Gelatin is made up of protein molecules in water. (If you're not sure what molecules are, read "Resource Unit 24: Atoms, Molecules, and Ions.") Although the protein molecules are too small to be seen, they are very long and slender. The molecules are all tangled up with each other. Between the tangles, in the empty spaces, are water molecules. Look at Figure 8-1 below.

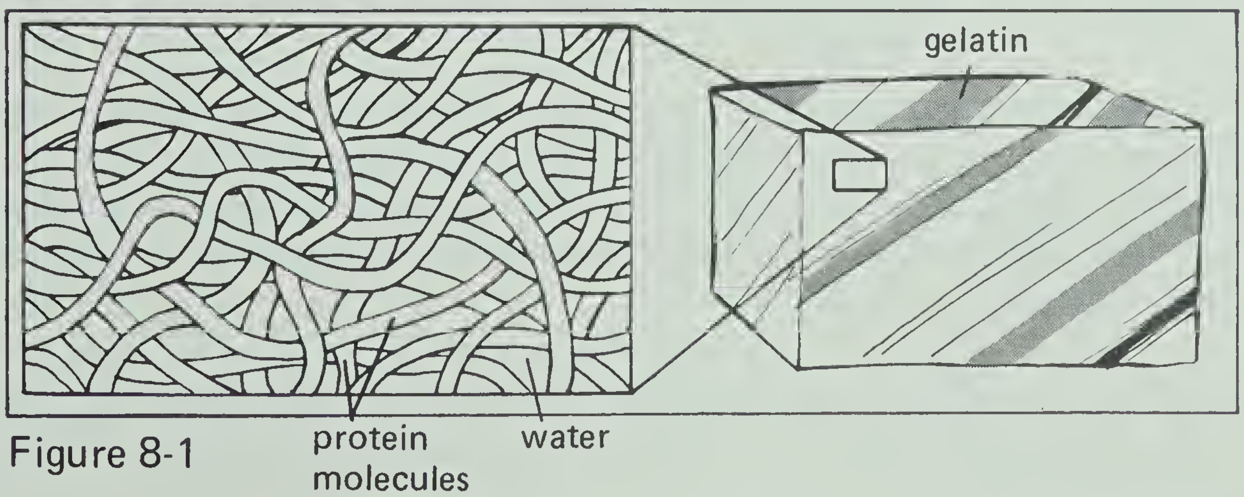
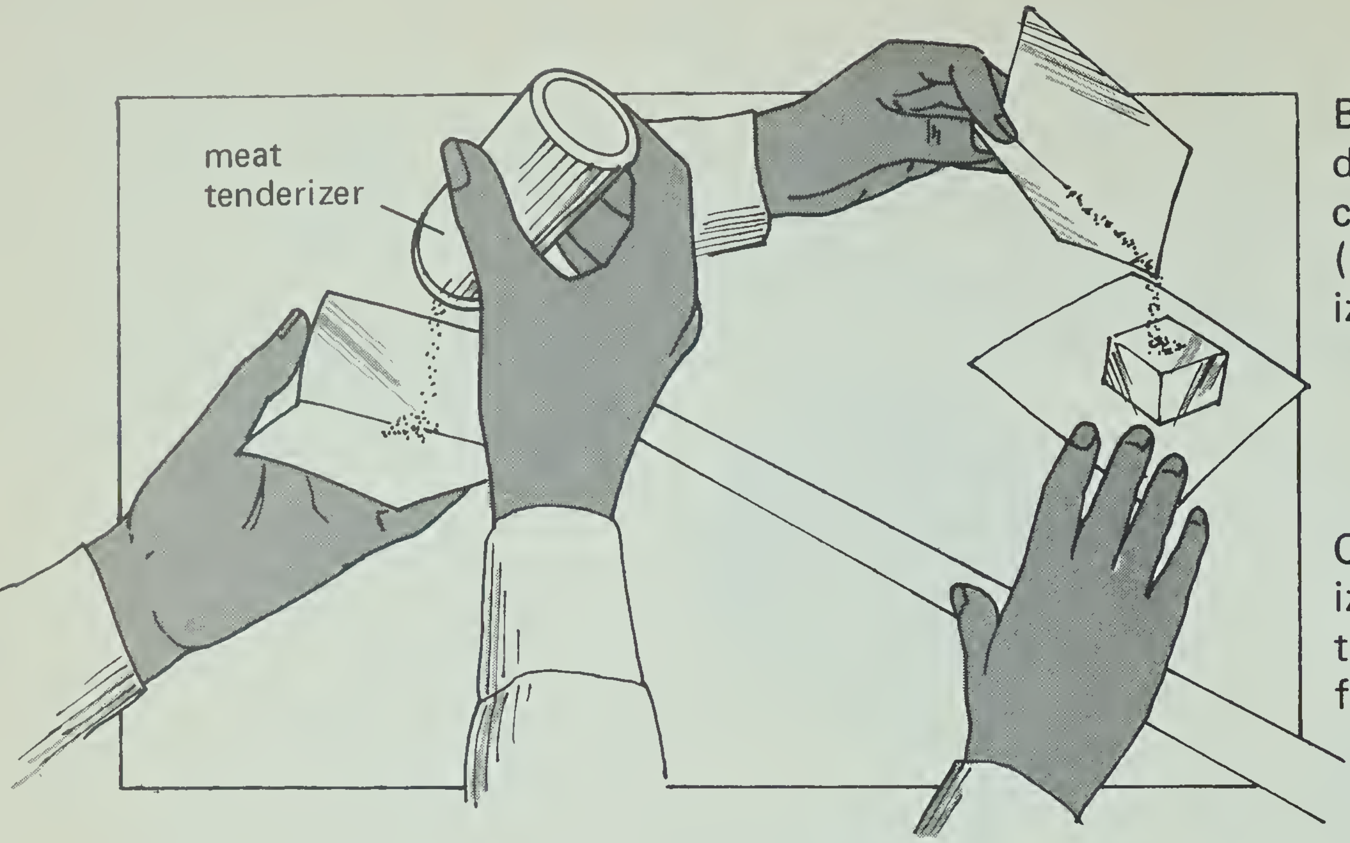


Figure 8-1



B. Fold the piece of paper down the middle, and make a crease. Sprinkle about 1 ml (¼ teaspoon) of meat tenderizer onto the paper.

C. Pour all of the meat tenderizer onto one spot on top of the gelatin. Watch the gelatin for a few minutes.

8-2. The gelatin looked wet around the meat tenderizer.

● 8-2. Describe what you observed when you put the meat tenderizer on the gelatin.

A chemical reaction occurred on the gelatin. Meat tenderizer contains a type of chemical catalyst known as an *enzyme*. (A catalyst is a substance that starts a chemical reaction.) This particular enzyme, called *papain*, breaks up the threadlike molecules of protein. This untangles the protein molecules and frees the water that was trapped in the tangles. What you saw when you put the meat tenderizer on the gelatin was that freed water. Look at Figure 8-2 below.

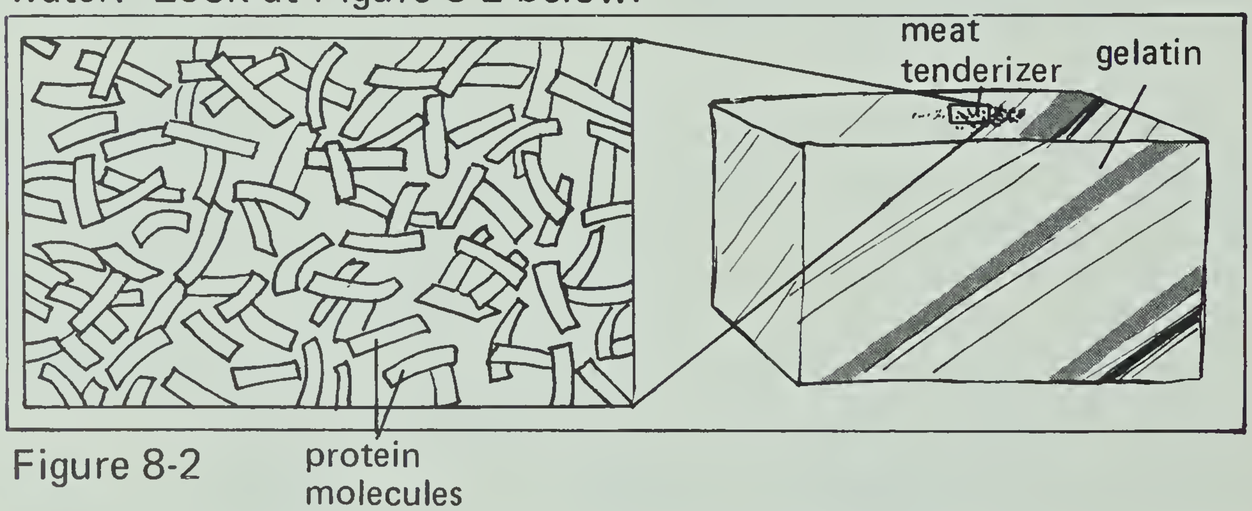


Figure 8-2

Raw pineapple contains an enzyme called *bromalin* that does the same thing as papain. That's why a gelatin dessert made with raw pineapple won't "set." When you eat raw pineapple, the enzyme breaks up some of the protein molecules in your tongue and cheeks.

8-3. Pineapple contains an enzyme that breaks up protein molecules in your mouth and tongue.

● 8-3. If you ate a lot of raw pineapple, why would your mouth and tongue feel tender?

You'd have to eat a lot of raw pineapple to cause any permanent damage to your mouth and tongue. And there's no problem with eating cooked pineapple or cooked meat with tenderizer on it, because cooking destroys these enzymes. But uncooked meat tenderizer contains the enzyme in concentrated form. Eating even a little of it could damage your mouth and tongue.

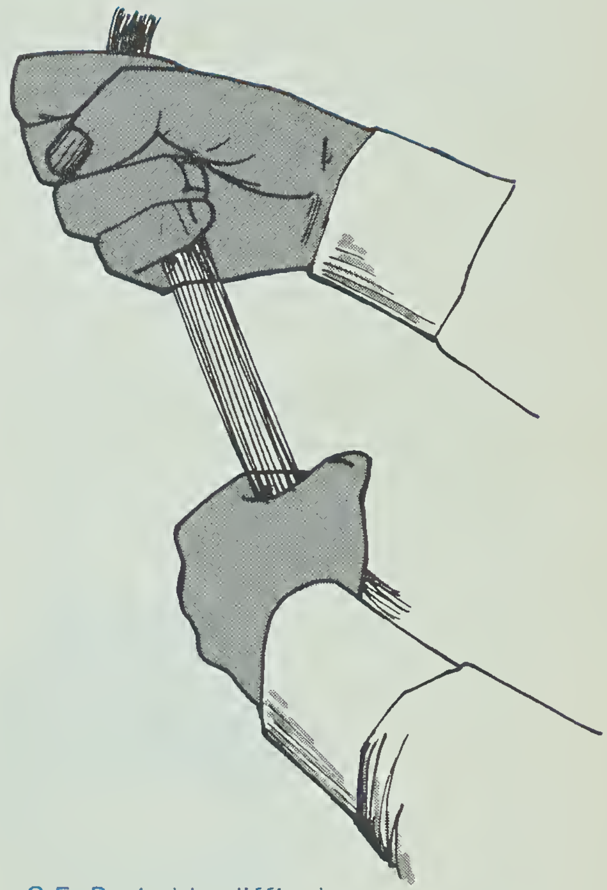
☆ **8-4. What is the active ingredient in meat tenderizers?**

8-4. An enzyme called *papain*

To understand better how chemists think enzymes such as the one in meat tenderizers work, look at Figure 8-3 below. It shows some threads lined up more or less parallel. The threads represent protein molecules in a piece of meat, in your cheek or tongue, or in some food in your stomach.

Figure 8-3

Figure 8-4



• 8-5. Suppose you held all the threads in Figure 8-3 above at both ends and tried to pull them apart. Would it be easy or difficult?

8-5. Probably difficult

• 8-6. If the threads were cut in different places, like those in Figure 8-4 above, would it be easier or more difficult to pull them apart?

8-6. Easier

Meat tenderizers soothe stings from coral, jellyfish, sea urchins, etc., and mosquito and fly bites.

When many of the threads are broken, they can be drawn apart more easily. Meat tenderizer breaks up many of the protein molecules in tough meat, making it easier to chew.

The papain in meat tenderizer comes from the papaya fruit, which is grown mostly for that purpose. Just as with pineapple, eating a little raw papaya probably won't cause any harm to your mouth. But eating a lot could. Fortunately, heating destroys the enzyme. So, you can eat tenderized meat after it has been cooked without tenderizing your mouth.

☆ **8-7. How do meat tenderizers work?**

8-7. They break up the long protein molecules that make meat tough.

ACTIVITY EMPHASIS: Energy changes involved in dissolving and crystallizing obey the law of conservation of energy: No energy is created or destroyed; the distribution of motion energy in the system changes.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter.

ACTIVITY 9: ENERGY AND CRYSTALS

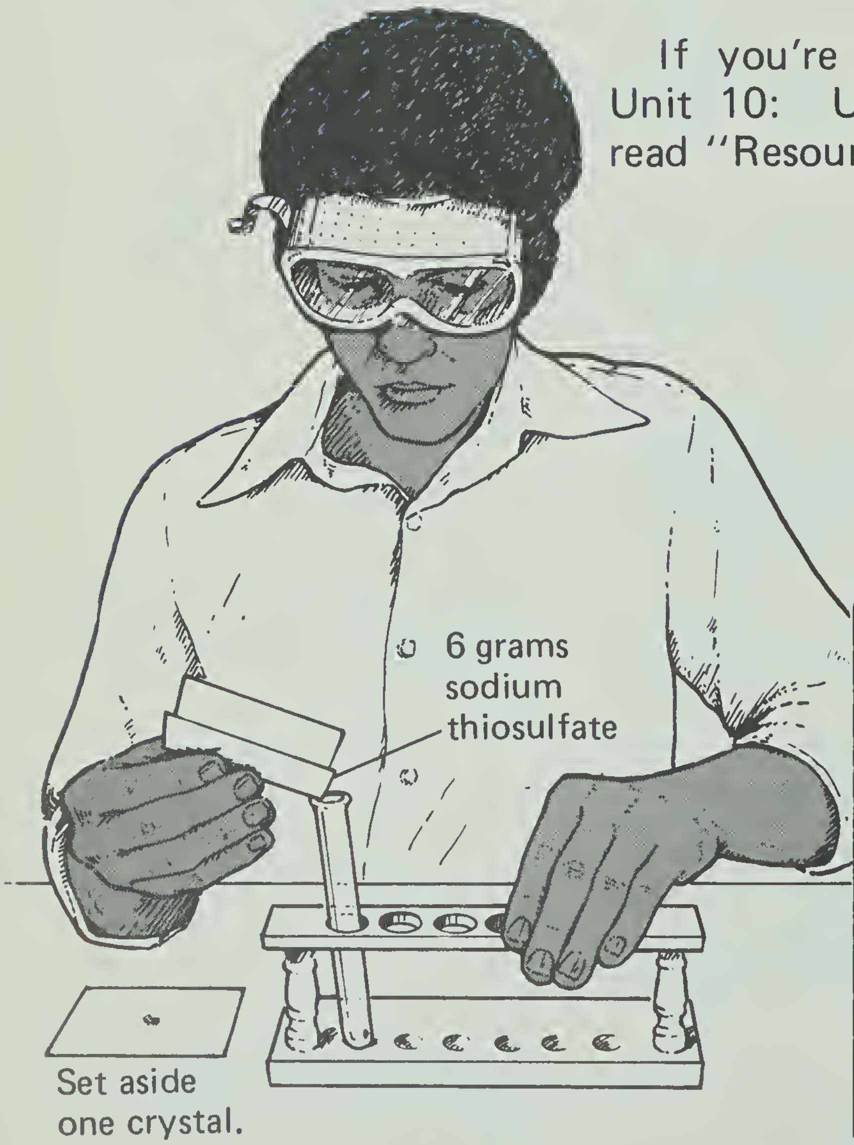
One of the basic principles of science is that in ordinary (nonnuclear) changes, energy is neither created nor destroyed. The *amount* of energy after the change is the same as before the change. This idea is called the *law of conservation of energy.*

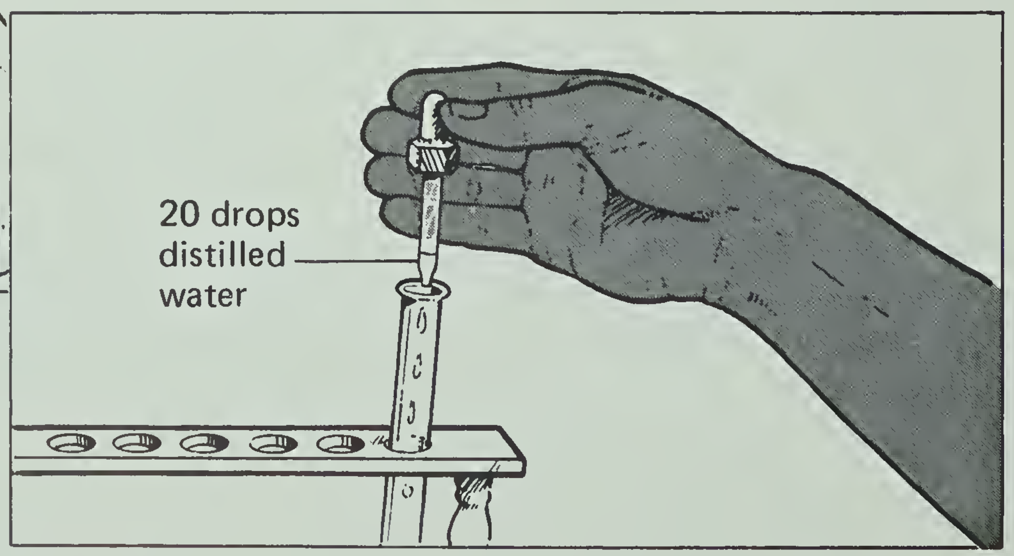
You can get a better idea of this by investigating what happens when crystals dissolve and then re-form from solution. You'll need about twenty minutes and the following materials.

balance
distilled water in dropping bottle
250-ml beaker two-thirds full of tap water
sodium thiosulfate, $Na_2S_2O_3 \cdot 5\,H_2O$
medium test tube with stopper
test-tube holder
Bunsen burner or other heat source
watch or clock with second hand
paper towel
forceps
test-tube rack
safety goggles

If you're not sure how to use a balance, read "Resource Unit 10: Using Balances." If you need help using a burner, read "Resource Unit 17: Using a Burner." Then begin Step A.

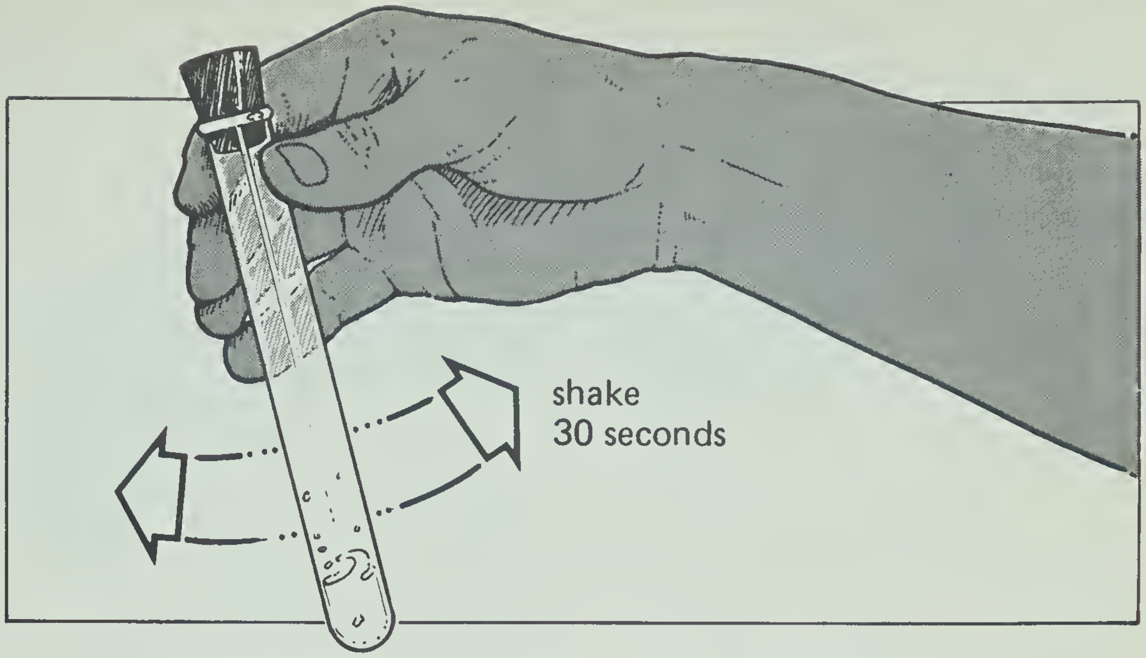
A. Weigh out 6 grams of sodium thiosulfate. Put all the crystals except one into the test tube. Keep the one extra crystal aside to use later. Add 20 drops of distilled water to the test tube.





B. Stopper the test tube, and shake it back and forth for 30 seconds. The crystals will begin to dissolve.

C. Feel the bottom of the test tube.



● 9-1. What happens to the temperature of the mixture in the test tube as the crystals dissolve?

9-1. It decreases.

Sodium thiosulfate crystals are made up of sodium ions, Na^+, and thiosulfate ions, $S_2O_3{}^{2-}$. (See "Resource Unit 24: Atoms, Molecules, and Ions" if you don't know what ions are.) These + and – ions are arranged opposite each other in all directions to form each crystal.

In order for the crystals to dissolve, the ions must be freed to mix with the water. It takes energy to free the ions from each other. One way the ions can get this energy is from the water molecules. As the ions get energy, the crystals dissolve, and the temperature of the solution decreases. Look at Figure 9-1 below.

Of course, there are water molecules as well as the two different sizes of ions in the sodium thiosulfate crystal lattice. In addition, several water molecules "add energy" to dissolving ions by hydrating them. Figure 9-1 here and Figures 9-2 (page 42) and 9-4 (page 45) omit these details in order to focus simply on the energy-transfer process.

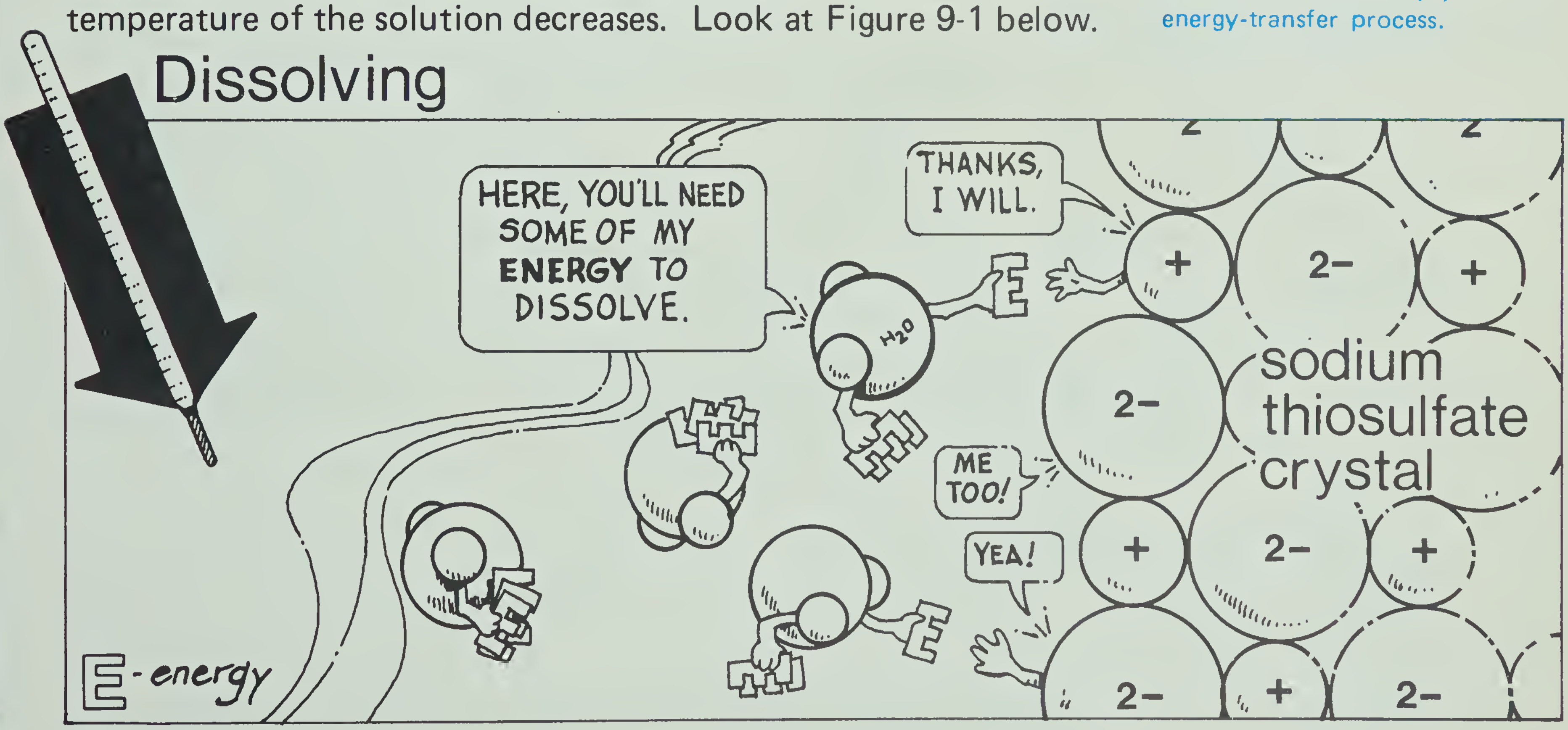
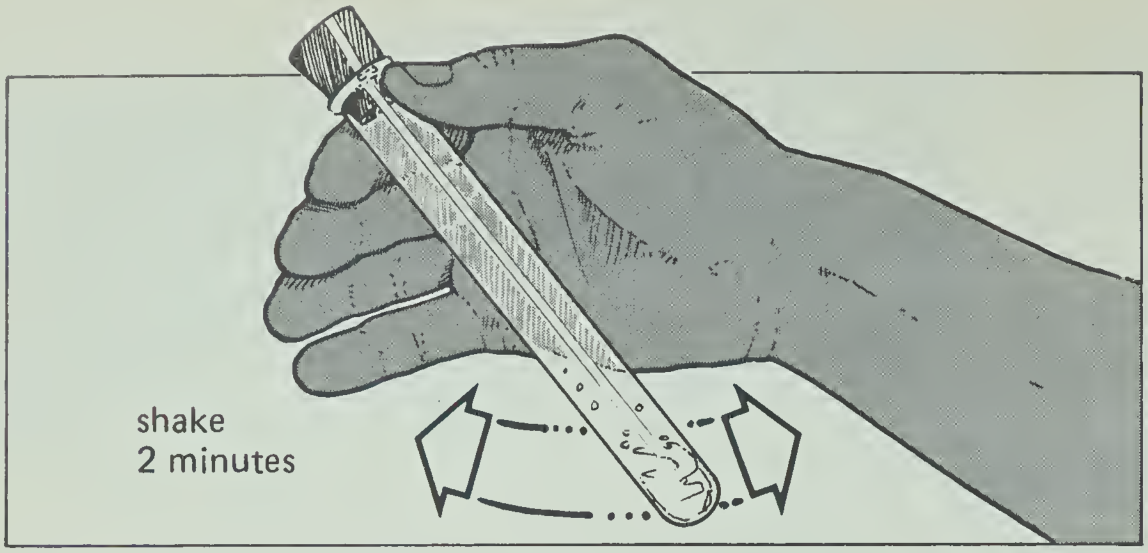


Figure 9-1

● 9-2. Why does the temperature decrease as the crystals begin to dissolve?

9-2. The sodium and thiosulfate ions get energy from the water molecules as they dissolve.



D. Shake the test tube for two more minutes. This will add some more energy to the system.

9-3. No; some solid still remains in the test tube.

• 9-3. Does all the sodium thiosulfate dissolve? How do you know?

At room temperature, not all the ions can get enough energy from the water to dissolve. There are probably about four grams of sodium thiosulfate not dissolved in your test tube.

9-4. About two grams

• 9-4. About how many grams of sodium thiosulfate are dissolved in your saturated solution?

9-5. Add more energy to the ions.

• 9-5. Look at Figure 9-2 below. How can the rest of the crystals be dissolved?

Saturated Solution

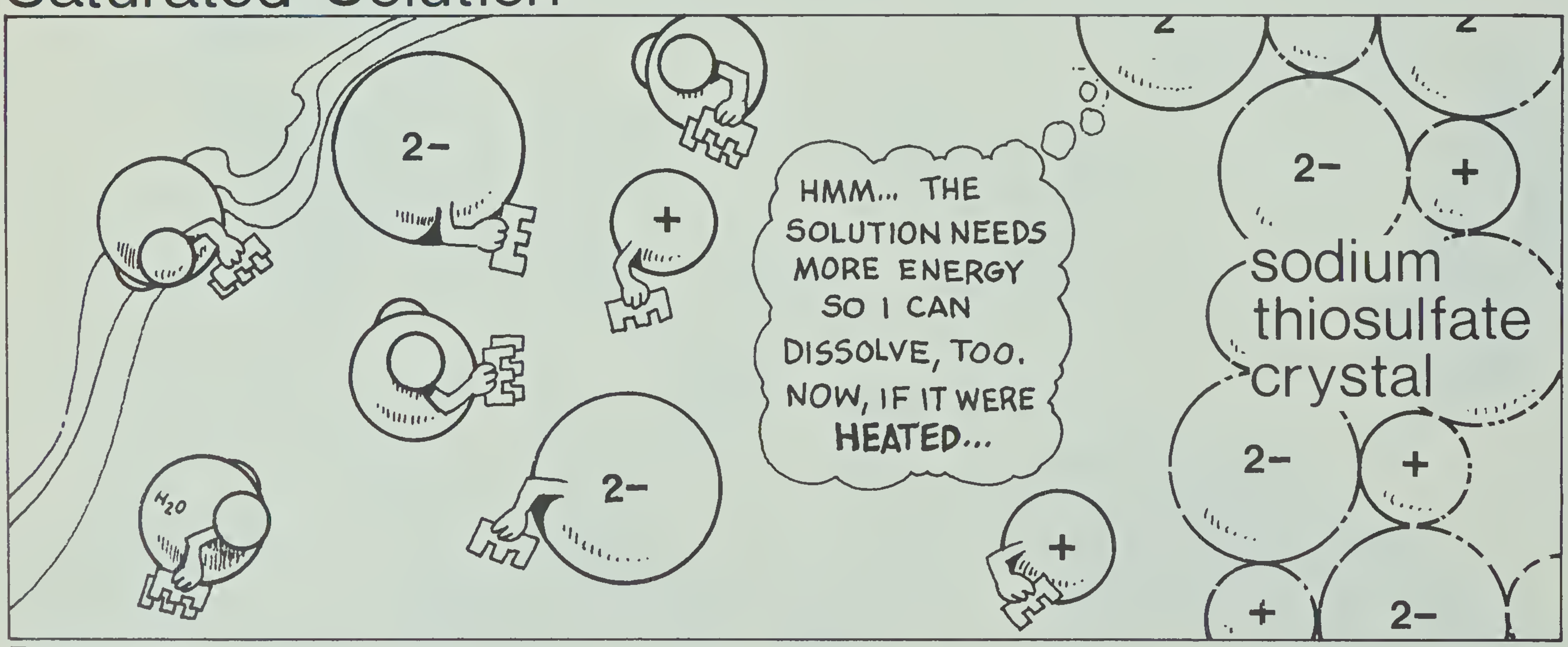


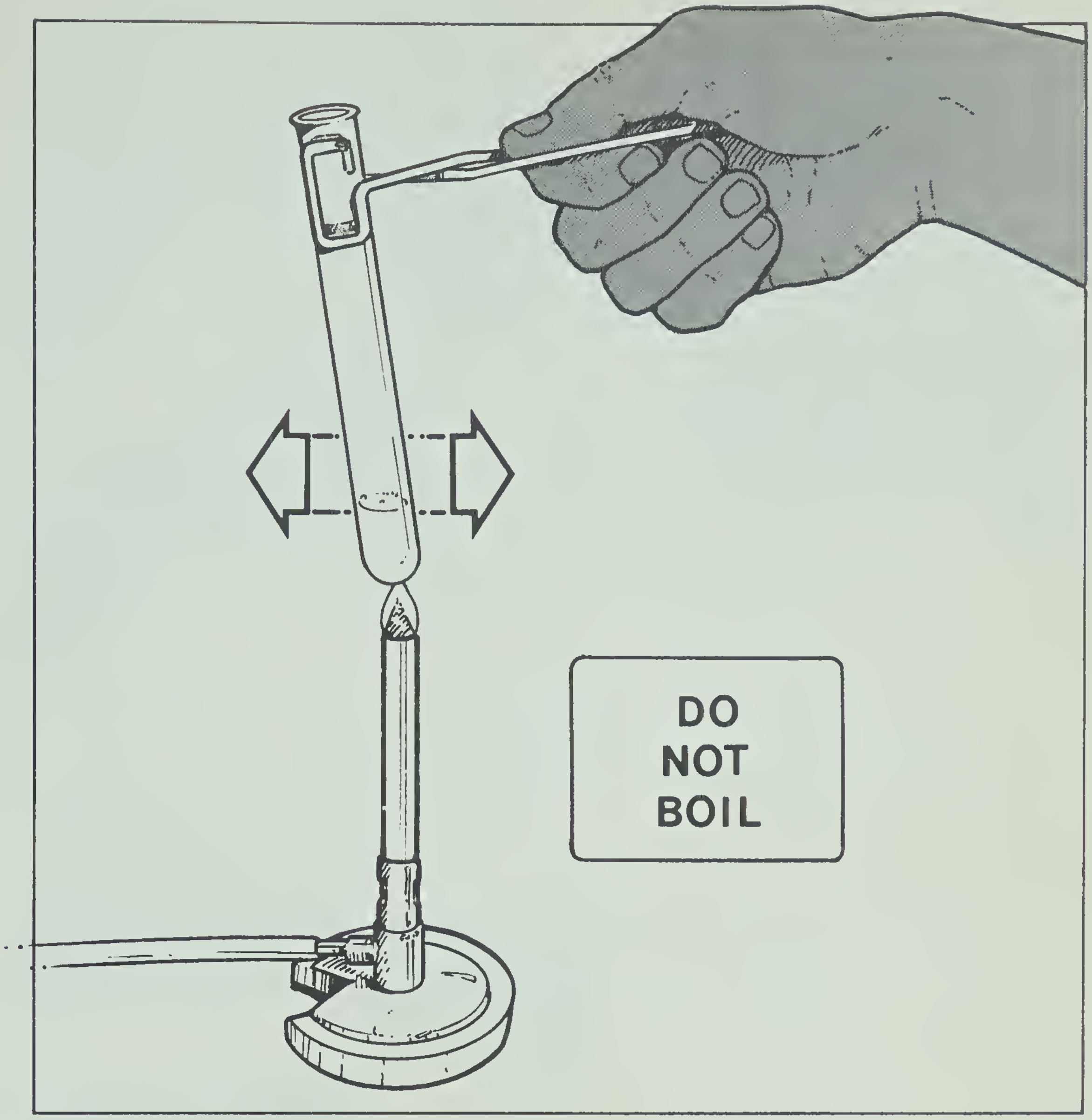
Figure 9-2

The only way to add more energy is to heat the mixture.

CAUTION	Do not heat the test tube with the stopper still in it.

E. Light the burner. Remove the stopper from the test tube. Using the test-tube holder, move the test tube back and forth above the upper part of the flame. Keep the mouth of the test tube pointed away from yourself and others.

F. As soon as all the crystals have dissolved, remove the test tube from the flame and turn off the burner. *Do not let the solution boil.*



You now have a hot, nearly saturated solution of sodium thiosulfate. If you cool this solution carefully, it will become supersaturated.

G. Put the test tube into the beaker of water and allow it to stand for 5 minutes. Do not touch, move, or disturb the test tube. Let it remain absolutely undisturbed.

9-6. Energy is leaving the test tube and solution.

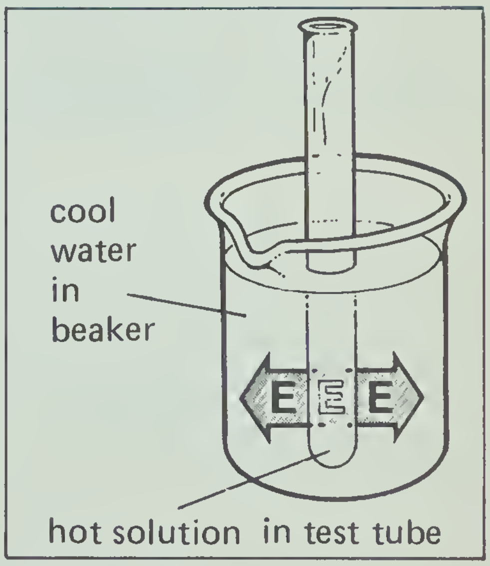
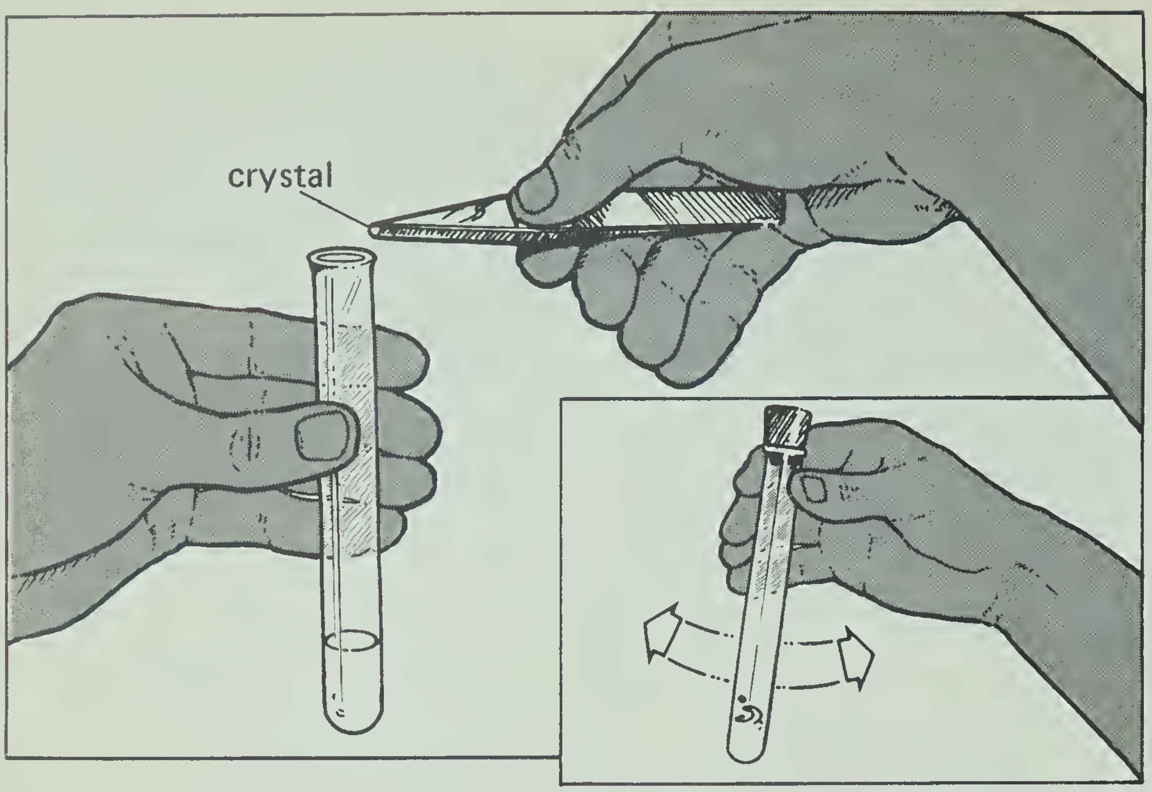


Figure 9-3

● 9-6. What is happening in Step G? (Hint: Look at Figure 9-3.)

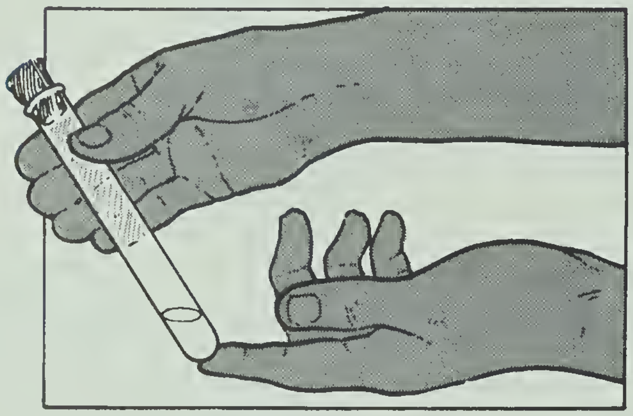
A supersaturated solution won't stay that way. The extra dissolved ions will recrystallize out of solution.



H. Take the test tube out of the water. Wipe the outside of the test tube with the paper towel. Using the forceps, add the crystal you put aside in Step A. Then replace the stopper. Shake the test tube back and forth several times.

9-7. The extra dissolved sodium thiosulfate recrystallized out of solution.

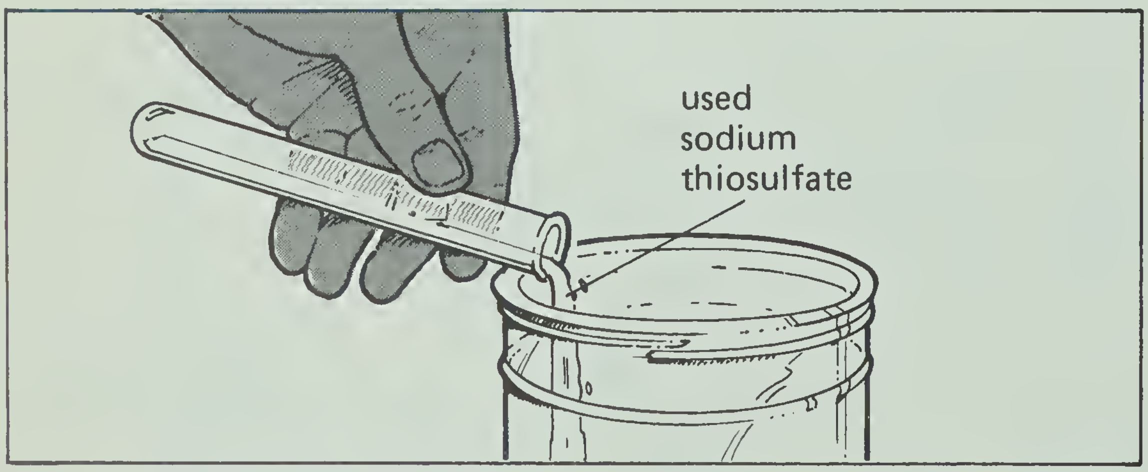
● 9-7. Describe what happened when you dropped the crystal into the supersaturated solution.



I. Feel the bottom of the test tube with your fingers.

9-8. It felt warm.

● 9-8. What did you notice about the test tube when you felt it?



J. Dispose of the sodium thiosulfate as directed by your teacher.

If you want to recover the sodium thiosulfate, see "Background Information" (page ATE 9) in the ATE front matter.

Maybe you were surprised to find that the temperature of the solution increased in Step I. But remember that the energy used to dissolve the ions wasn't destroyed. It was still in the solution, in the motion of the dissolved ions. So when the ions re-formed crystals, the energy was given back to the water molecules. You felt the energy as heat in Step I.

Crystallizing from Supersaturated Solution

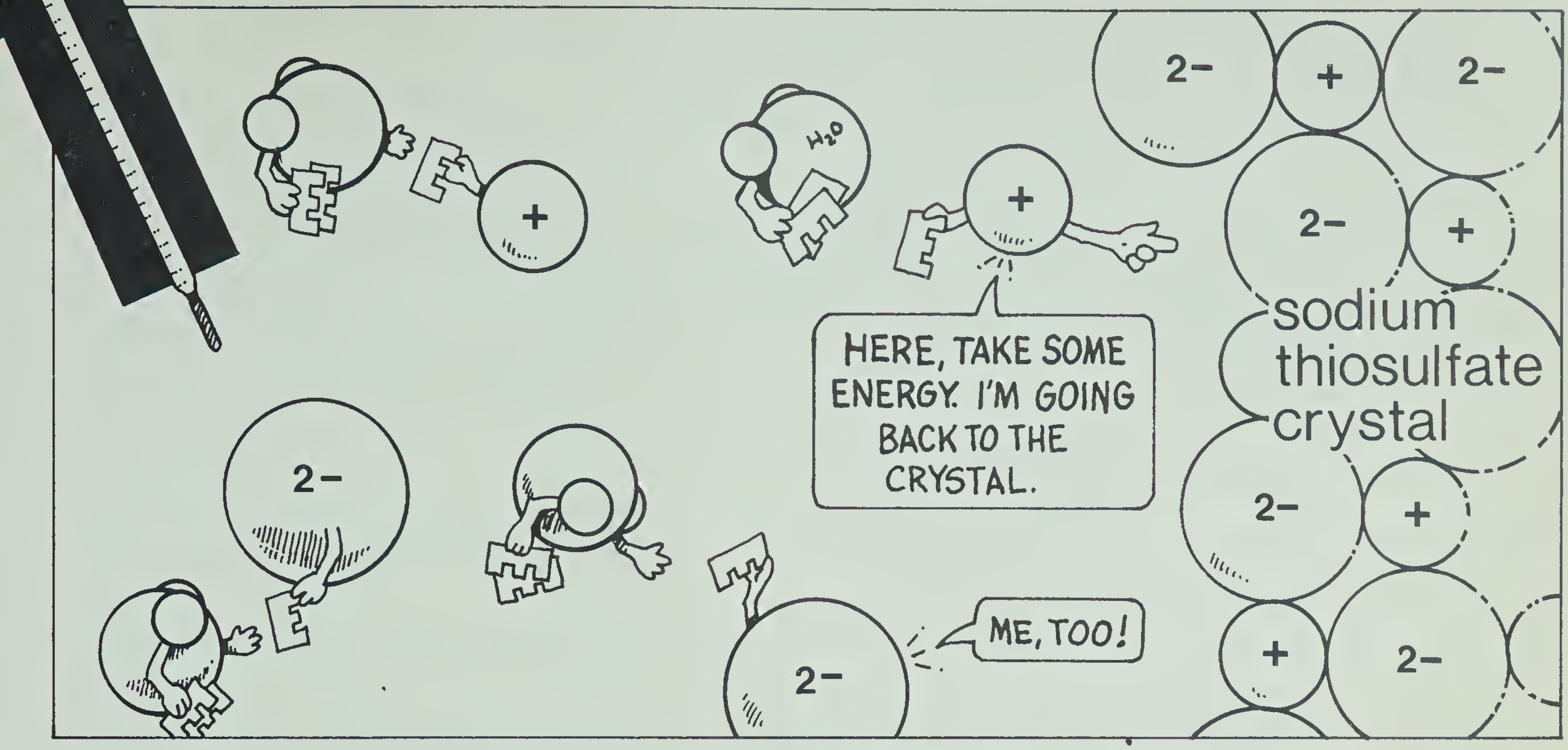


Figure 9-4.

• 9-9. Why does the temperature increase as sodium thiosulfate recrystallizes from a supersaturated solution?

9-9. The sodium and thiosulfate ions release energy as they re-form the crystals.

In your investigation, no energy was destroyed or created. Some energy was given to the dissolved ions for a while, and then it was given back. That's conservation of energy.

☆ **9-10. Describe, in terms of the law of conservation of energy, the energy changes involved when sodium thiosulfate dissolves in water and then recrystallizes.**

9-10. Energy is absorbed when the ions in sodium thiosulfate crystals dissolve and released when the crystals re-form from solution. No energy is created or destroyed in either process.

If this is the first time you've seen the law of conservation of energy, don't worry if you aren't sure about it. You'll see the idea in other minicourses. If you have seen it before and you want extra help, you might look at "Resource Unit 18: Matter and Energy."

ACTIVITY 10: PLANNING

If you plan to do Activities 11 and 12, do Activity 11 first.

Activity 11 — Page 47

Objective 11-1: Define *acid* and *base* in terms of hydrogen ions.

Sample Question: Match the substance with its definition.

Substance	Definition
A. Acid	1. releases hydrogen ions
B. Base	2. shares hydrogen ions
	3. takes hydrogen ions

Objective 11-2: Explain the difference between strong acids and weak acids.

Sample Question: Tell whether the following statement is true or false. A strong acid releases more hydrogen ions than a weak acid does.

Activity 12 — Page 52

Objective 12-1: Tell why the solutions of some salts, such as sodium carbonate, sodium hydrogen carbonate, and ammonium chloride, in water are acidic or basic.

Sample Question: When washing soda (sodium carbonate) dissolves in water, it forms

A. an acidic solution because sodium ions take hydrogen ions.
B. a basic solution because carbonate ions take hydrogen ions.
C. an acidic solution because sodium ions release hydrogen ions.
D. a basic solution because carbonate ions release hydrogen ions.

Objective 12-2: Describe how water can act as either an acid or a base.

Sample Question: Water acts as an acid when

A. it releases hydrogen ions to other water molecules.
B. it takes hydrogen ions from certain ions.
C. it releases hydrogen ions to certain ions.

Activity 13 — Page 56

Objective 13-1: Describe and explain the effects that dissolved substances have on the freezing and boiling points of water.

Sample Question: The more salt that is dissolved in a certain amount of water,

A. the lower the freezing point, because there are fewer particles in solution.
B. the lower the freezing point, because there are more particles in solution.
C. the higher the freezing point, because there are fewer particles in solution.
D. the higher the freezing point, because there are more particles in solution.

Answers: **11-1.** A1, B3; **11-2.** True; **12-1.** B; **12-2.** A, C; **13-1.** B

ACTIVITY 11: GIVING AND TAKING

In the core, you learned that acids and bases could be defined in terms of their effects on litmus paper. You saw that bases turn red litmus blue and acids turn blue litmus red. But this definition does not get at the idea of what acids and bases do that makes them different from other substances. In this activity, you'll learn a definition that does just that.

Before you begin, you might want to review information on formulas, equations, and ions. Look at "Resource Unit 22: Chemical Shorthand" and "Resource Unit 24: Atoms, Molecules, and Ions." Also, check the index of *Actions and Reactions*.

ACTIVITY EMPHASIS: Acids form hydrogen ions and bases take hydrogen ions. At equal solution concentrations, a strong acid releases more hydrogens as hydrogen ions than a weak acid does.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

- 11-1. What is an ion?

11-1. A particle that has gained or lost one or more electrons.

Acids are usually corrosive. A nail will rust faster in vinegar or lemon juice than it will in plain water or air. Some acids, such as sulfuric acid and hydrochloric acid, are more corrosive than other acids. Most of the very corrosive acids are called *strong acids*. Figure 11-1 below lists most of the strong acids.

Of course, salts and bases also have corrosive effects.

STRONG ACID	FORMULA
Hydrochloric acid	HCl
Sulfuric acid	H_2SO_4
Nitric acid	HNO_3
Hydriodic acid	HI
Hydrobromic acid	HBr
Perchloric acid	$HClO_4$

Figure 11-1

All of the other acids that you are likely to hear about are called *weak acids*.

- 11-2. Is acetic acid (the acid in vinegar) a strong or a weak acid?

11-2. A weak acid

An acid is present if hydrogen ions are formed when a substance is dissolved in water. The hydrogen ions, which are positive, are associated with water molecules. For example, when perchloric acid is dissolved in water, it forms positive hydrogen ions and negative perchlorate ions.

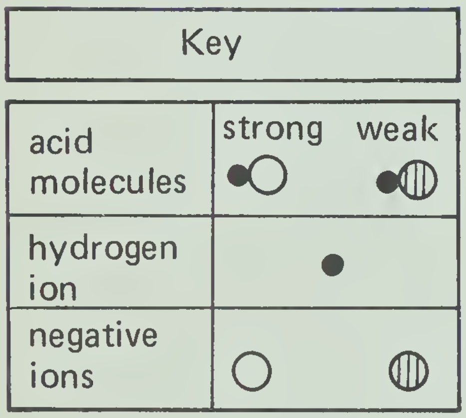
$$HClO_4 \xrightarrow[\text{dissolved in water}]{} H^+_{(aq)} + ClO_4^-{}_{(aq)}$$

The symbol *aq*, as in $H^+_{(aq)}$, stands for *aqueous*. It means that the ions are associated with water molecules.

11-3. $HCl \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$ dissolved in water

● 11-3. Write the equation for hydrochloric acid dissolving in water.

Hydrogen ions are also formed when a weak acid, such as acetic acid, is dissolved in water. But not all of the dissolved acetic acid molecules release their hydrogens as hydrogen ions. Strong acids release their hydrogens much more readily than weak acids do. Look at Figure 11-2 below.



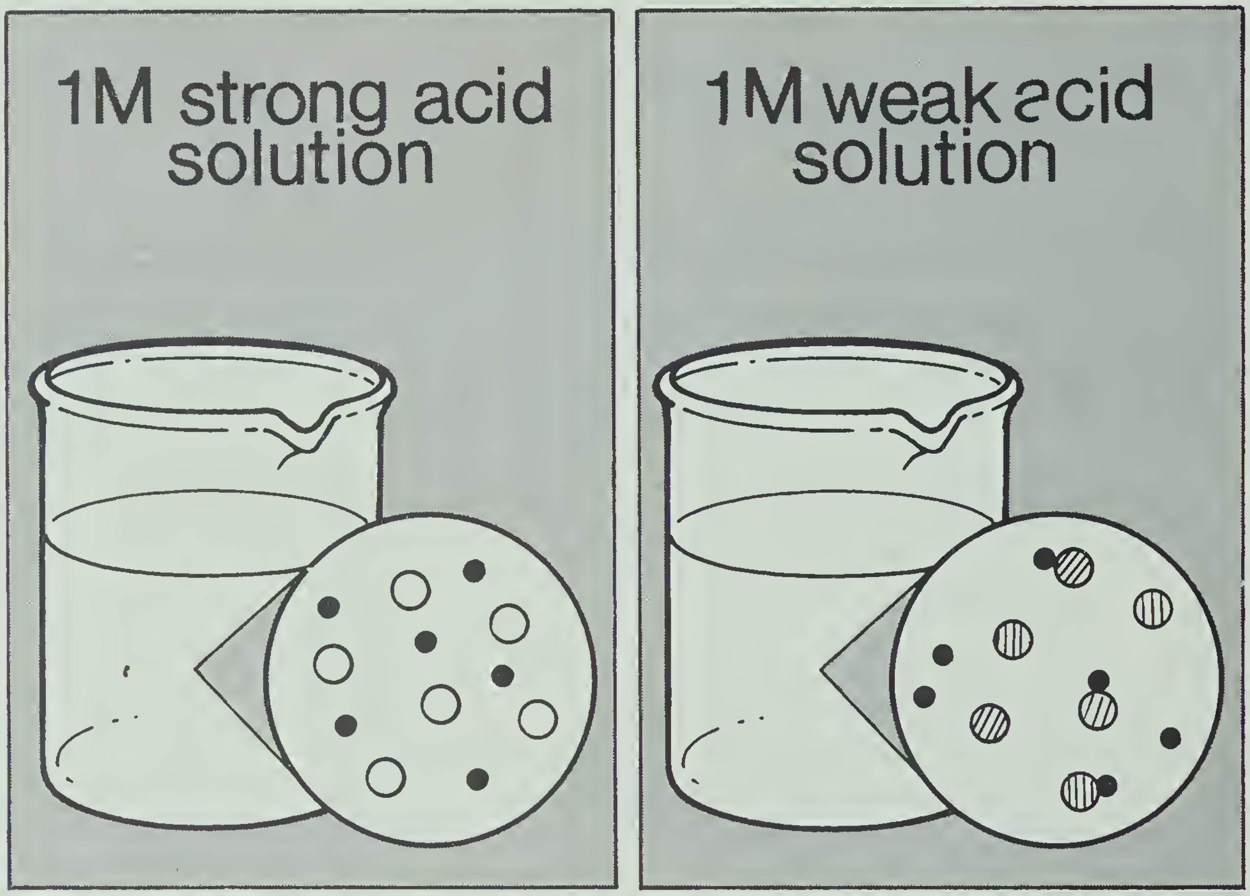


Figure 11-2

Remember that all of the dissolved ions and molecules are associated with water molecules.

11-4. There are more hydrogen ions in the strong acid solution; the strong acid releases hydrogens more readily than the weak acid does.

☆ **11-4. How does a solution of a strong acid differ from a solution of a weak acid of the same concentration? Why?**

It is the hydrogen ions that make blue litmus turn red and that make acids corrosive. Thus, many chemists prefer to define *acids* as "substances that form hydrogen ions" instead of "substances that turn blue litmus red."

11-5. Hydrogen ions

☆ **11-5. What kind of ions do all acids form when they are dissolved in water?**

Bases are sort of the opposite of acids. Bases are defined as "substances that take hydrogen ions." To check out some acids and bases, do this investigation. You'll need about fifteen minutes and the following materials.

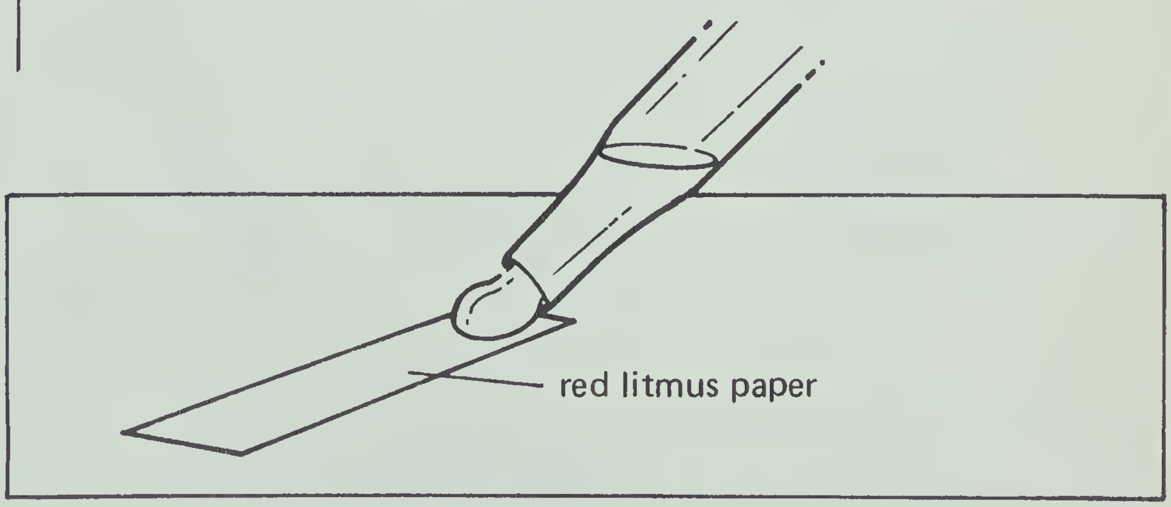
- 3 red litmus paper strips
- 3 blue litmus paper strips
- sodium hydroxide, NaOH, solution, 0.1M, in dropping bottle
- calcium hydroxide, $Ca(OH)_2$, solution (saturated) in dropping bottle
- potassium hydroxide, KOH, solution, 0.1M, in dropping bottle
- barium hydroxide, $Ba(OH)_2$, solution, 0.1M, in dropping bottle
- hydrochloric acid, HCl, solution, 0.1M, in dropping bottle
- acetic acid, $HOOCCH_3$, solution, 0.1M, in dropping bottle
- safety goggles

CAUTION	Some of these solutions can be dangerous. If you spill any, wash immediately and thoroughly with lots of water. Tell your teacher. Wear your safety goggles.

If you're not sure how to do litmus-paper tests, review Activity 2 for the instructions. Then continue with this investigation.

A. In your notebook, make a table like this one. Make it large enough for all six solutions you will test. If you have the results for some of them from Activity 2, you don't need to test them again. Just enter the results in the table in your notebook.

SOLUTION	TURNS BLUE LITMUS RED	TURNS RED LITMUS BLUE	IS AN ACID	IS A BASE	RELEASES HYDROGEN IONS	TAKES HYDROGEN IONS
sodium hydroxide, NaOH						



B. Test each liquid with red litmus paper. Record your results in your table.

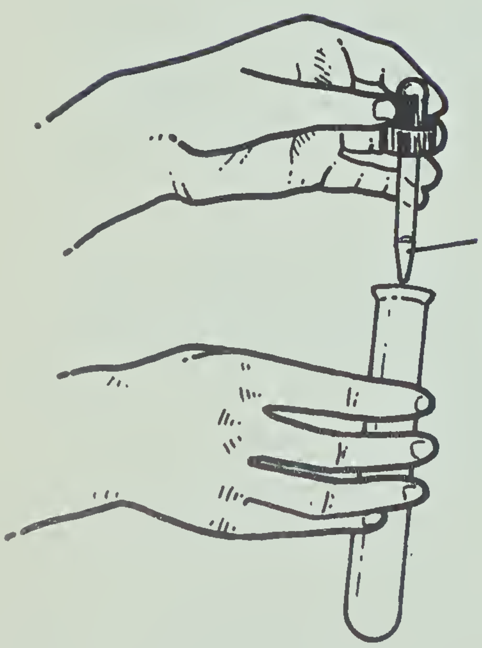
C. Test each solution with blue litmus paper. Record your results in your table.

11-6. *Hydroxide*

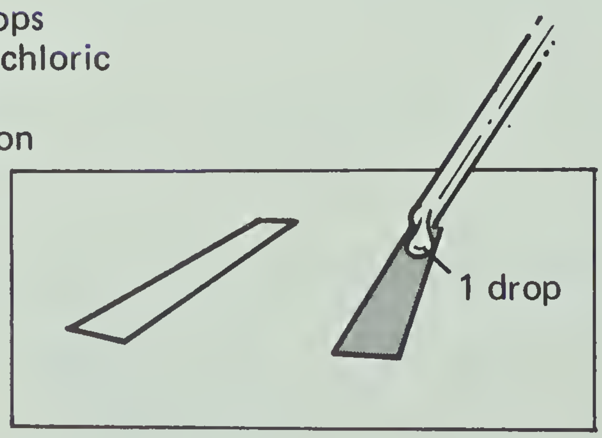
• 11-6. What word is common to the names of the substances that tested as basic?

The four basic solutions you tested all contain negative hydroxide ions, OH^-, associated with water molecules. From that, you could say that such substances release hydroxide ions. But actually, they also take hydrogen ions. To see this process in action, you'll need about twenty minutes and the following materials.

3 red litmus paper strips
3 blue litmus paper strips
sodium hydroxide, NaOH, solution, 0.1M, in dropping bottle
hydrochloric acid, HCl, solution, 0.1M, in dropping bottle
medium test tube
stirring rod
safety goggles



A. Put about 25 drops of the hydrochloric acid solution into the test tube.



B. Use the stirring rod to put 1 drop of the solution on the end of a red litmus strip and another drop on the end of a blue strip.

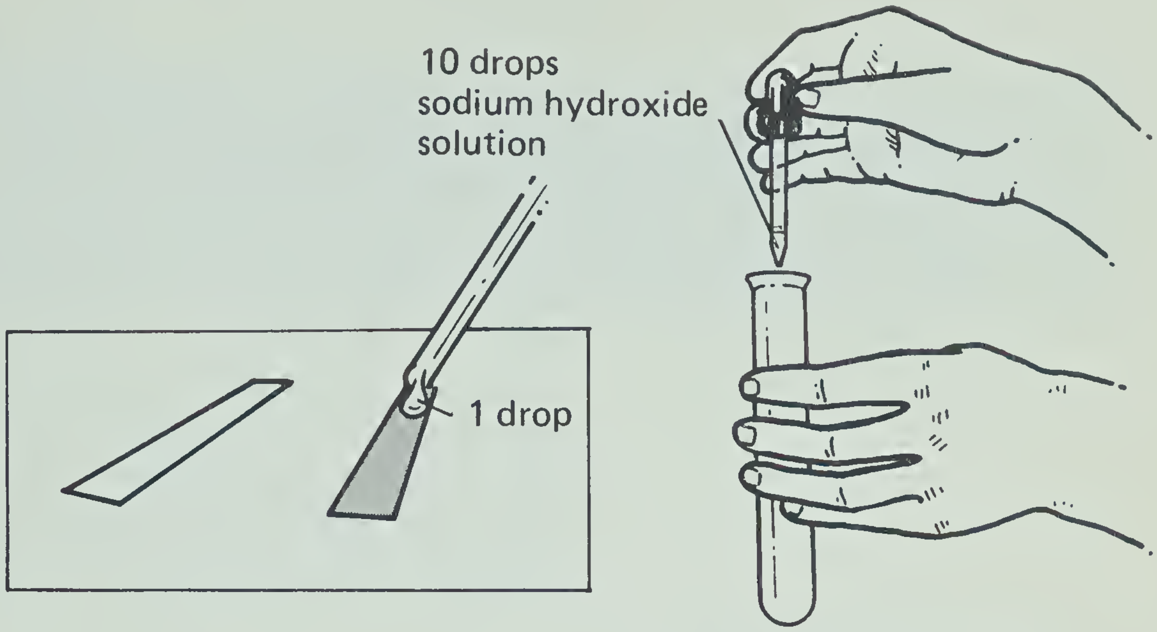
11-7. It turned red; it stayed red.

• 11-7. What happened to the blue litmus strip? The red litmus strip?

Since the blue litmus paper turned red, you know that there must be hydrogen ions in the hydrochloric acid solution.

C. Carefully measure exactly 10 drops of sodium hydroxide solution into the test tube containing the hydrochloric acid.

D. Stir the mixture. Then put one drop from the stirring rod onto a dry spot on each strip you used in Step B.

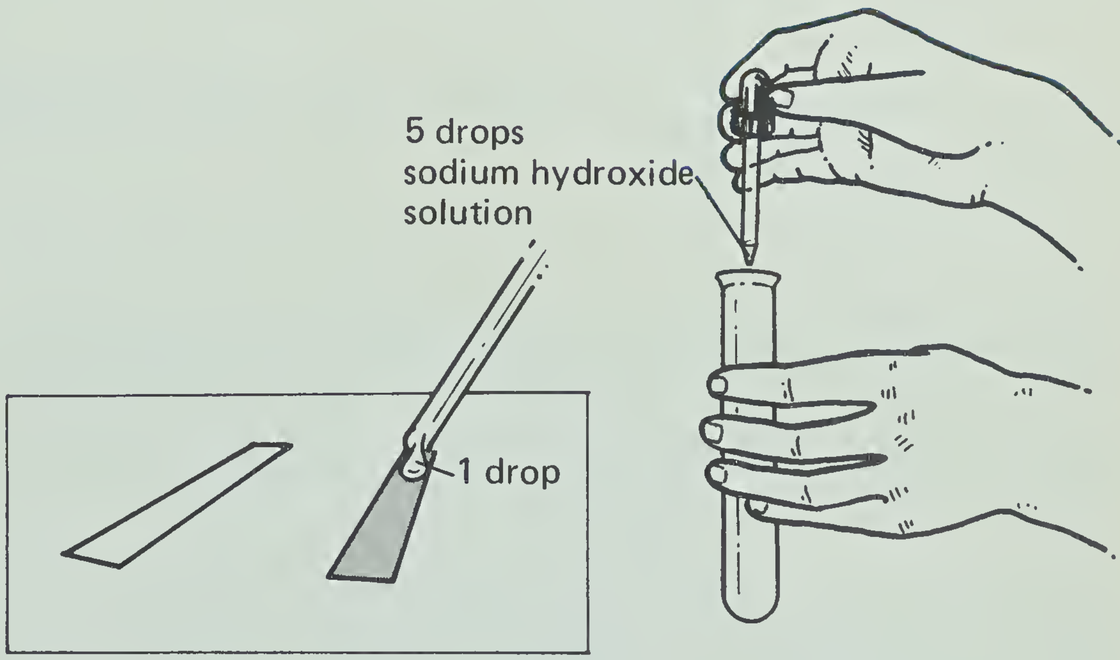


● 11-8. Are any hydrogen ions still present in the solution? How do you know?

11-8. Yes; the blue litmus turned red.

E. Add 5 more drops of the sodium hydroxide solution to the test tube and test the mixture, using new litmus strips.

F. Keep adding sodium hydroxide solution, 5 drops at a time, and testing the mixture until the blue litmus no longer turns red.



● 11-9. When the blue litmus no longer turns red, are there very many hydrogen ions in the solution?

11-9. No

Something in the sodium hydroxide took the hydrogen ions that were there to start with. If you repeated this investigation with calcium hydroxide, barium hydroxide, and potassium hydroxide, you'd find the same thing. After enough drops were added, the blue litmus would not turn red anymore.

● 11-10. What is in these solutions that can take hydrogen ions?

11-10. Hydroxide ions

Here is the equation for the reaction you observed between the hydrogen ions from the acid and the hydroxide ions from the base.

$$\underset{\text{From the acid}}{H^+_{(aq)}} + \underset{\text{From the base}}{OH^-_{(aq)}} \rightarrow H_2O \text{ or } H^+_{(aq)} + OH^-_{(aq)} \rightarrow HOH \text{ (or } H_2O)$$

Compounds such as sodium hydroxide and calcium hydroxide are often called *bases* because they release hydroxide ions in solution. But it's actually the hydroxide ions that are the bases — they take hydrogen ions, as all bases do. Any time hydroxide ions react with hydrogen ions, water molecules are formed.

11-11. They can take the hydrogen ions to form water molecules.

☆ **11-11. What can all bases do when hydrogen ions are present?**

ACTIVITY EMPHASIS: When some salts dissolve in water, ions that release hydrogen ions and ions that take hydrogen ions are released or formed. Hence, these solutions are acidic or basic. Water can act as both an acid and a base.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter.

ACTIVITY 12: SALTS AND WATER

When you hear the word *salt*, you probably think of table salt — sodium chloride, NaCl. But to a chemist, the word *salt* means that and a lot more too. A salt is an ionic compound — the product of a reaction of an acid with a base. (Check the index of *Actions and Reactions* for more about ionic compounds.)

Consider the reaction of solutions of hydrochloric acid and potassium hydroxide shown below.

$$\underset{\text{hydrochloric acid solution}}{H^+_{(aq)} + Cl^-_{(aq)}} + \underset{\text{potassium hydroxide solution}}{K^+_{(aq)} + OH^-_{(aq)}} \longrightarrow H_2O + K^+_{(aq)} + Cl^-_{(aq)}$$

When the water is evaporated from the solution, a solid ionic substance — potassium chloride, KCl — is left. KCl is a salt.

12-1. The positive ion came from the hydroxide compound; the negative ion came from the acid.

• 12-1. Did the positive ion in the KCl come from the acid or from the hydroxide compound? The negative ion?

Your answer to Question 12-1 on page 52 is usually true for salts — the positive ion comes from the hydroxide compound, the negative ion from the acid.

In addition, the positive ion is often a metal ion such as sodium, Na^+, or potassium, K^+. The negative ion is often either a nonmetal ion, such as chloride, Cl^-, or sulfide, S^{2-}, or an ion made up of a nonmetal and oxygen, such as nitrate, NO_3^-, or sulfate, SO_4^{2-}.

• 12-2. What acid and what hydroxide compound could react together to produce the salt sodium nitrate, $NaNO_3$?

12-2. Nitric acid, HNO_3, and sodium hydroxide, NaOH

Many salts form acidic or basic solutions when they dissolve in water. An investigation using three salts can help you understand this idea. It will take about fifteen minutes and these materials.

- ammonium chloride, NH_4Cl
- baking soda, $NaHCO_3$
- washing soda, Na_2CO_3
- distilled water
- 3 medium test tubes
- 3 soda straws
- test-tube rack
- strips of red and blue litmus paper
- grease pencil
- safety goggles

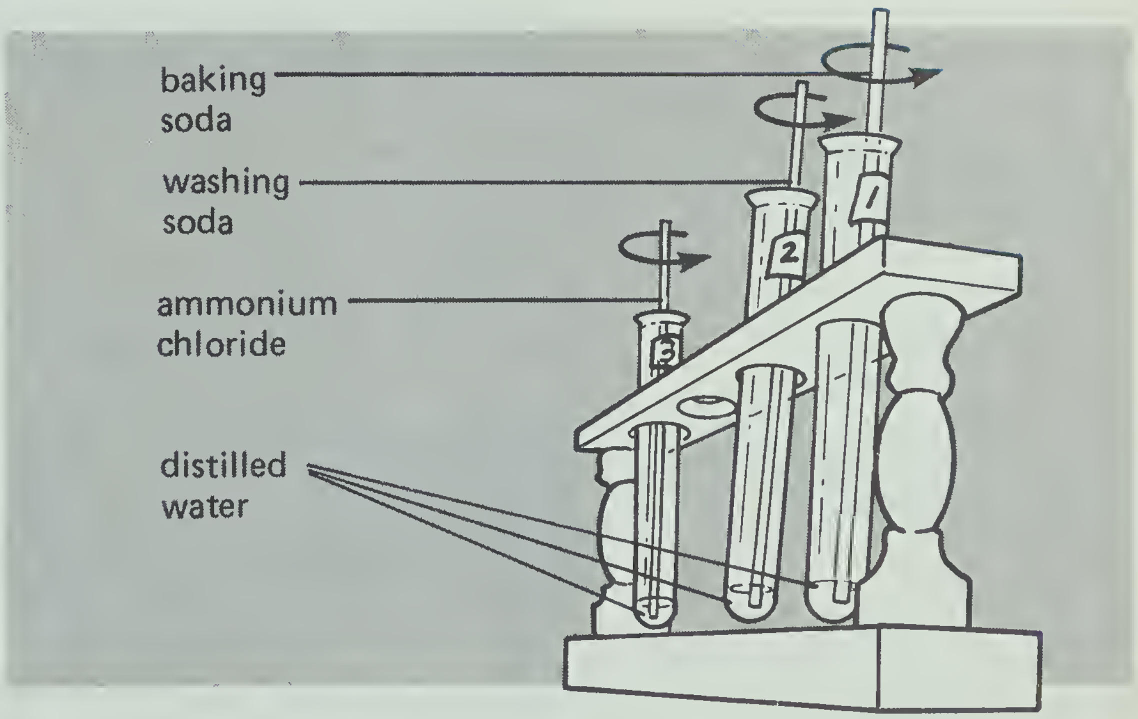
You may want students to use pH paper instead of litmus paper if it is available. Refer them to "Resource Unit 7: pH and How to Measure It."

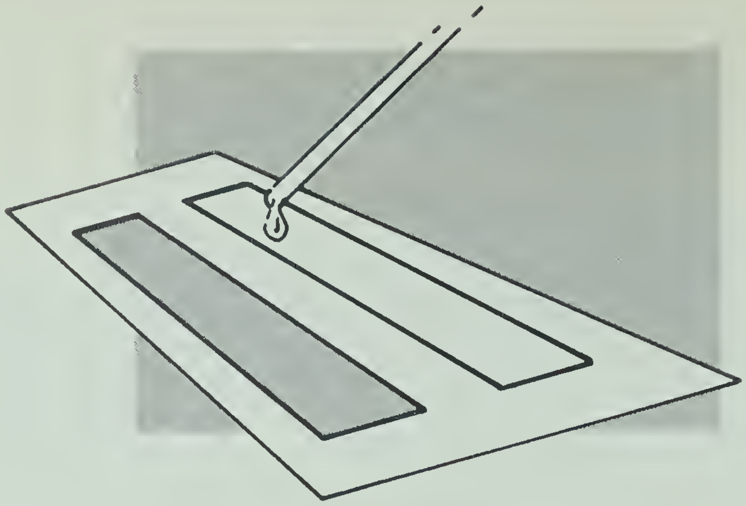
A. Label the test tubes *1, 2,* and *3.* Fill each test tube with distilled water to a depth of about 1 centimetre.

B. With one straw, put a little baking soda into the water in Test Tube 1. Leave the straw in the test tube.

C. Using a different straw for each, add washing soda to Test Tube 2 and ammonium chloride to Test Tube 3. Leave the straws in the test tubes.

D. Stir each mixture with its own straw until the solids have dissolved.





E. Use the straws to transfer a small drop of each solution to the strips of red and blue litmus paper.

12-3. Baking soda and washing soda form basic solutions. Ammonium chloride forms an acidic solution.

• 12-3. Does baking soda form an acidic or a basic solution? Washing soda? Ammonium chloride?

Substances such as $AlCl_3$ and $CaCl_2$ make acidic solutions by causing $H^+_{(aq)}$ to be formed.

Some salts are neither acidic nor basic in solution. Sodium chloride, for example, in solution forms $Na^+_{(aq)}$ and $Cl^-_{(aq)}$ ions, neither of which releases or takes hydrogen ions.

You found that baking soda and washing soda both form basic solutions and that ammonium chloride forms a slightly acidic solution. This means that there must be ions that take or release hydrogen ions in these three solutions. You can understand this process better by looking at the chemical equations involved.

The formula for baking soda is $NaHCO_3$ – it contains no hydroxide ions. In solution, sodium ions, $Na^+_{(aq)}$, and hydrogen carbonate ions, $HCO_3{}^-_{(aq)}$, are present.

$$NaHCO_{3\,(solid)} \xrightarrow[\text{dissolves}]{} Na^+_{(aq)} + HCO_3{}^-_{(aq)}$$

Some students may ask if $HCO_3{}^-_{(aq)}$ can also be an acid, according to this reaction:

$$HCO_3{}^-_{(aq)} \rightarrow H^+_{(aq)} + CO_3{}^{2-}_{(aq)}.$$

This reaction does occur, but to such a small extent that the $NaHCO_3$ solution is basic.

The sodium ions don't release or take hydrogen ions, so it's the hydrogen carbonate ions that take the hydrogen ions. They do this in two ways.

1. They take hydrogen ions from the water molecules that the hydrogen ions were associated with.

$$H^+_{(aq)} + HCO_3{}^-_{(aq)} \longrightarrow H_2CO_{3\,(aq)}$$

2. They take hydrogen ions from the water molecules themselves. (Note that H_2O is written as HOH.)

$$HOH + HCO_3{}^-_{(aq)} \rightarrow OH^-_{(aq)} + H_2CO_3$$

Thus, the hydrogen carbonate ion is a base. It will take a hydrogen ion from either a hydrogen ion associated with water molecules, $H^+_{(aq)}$, or from a water molecule, H_2O – whichever is handier.

☆ **12-4. Explain how a basic solution is produced when baking soda, $NaHCO_3$, dissolves in water.**

12-4. Hydrogen carbonate ions form when baking soda dissolves in water. These ions take hydrogen ions from water molecules and from $H^+_{(aq)}$.

Washing soda is sodium carbonate, Na_2CO_3. When it dissolves, sodium ions, $Na^+_{(aq)}$, and carbonate ions, $CO_3^{2-}{}_{(aq)}$, are formed.

$$CO_3^{2-}{}_{(aq)} + HOH \rightarrow HCO_3^-{}_{(aq)} + OH^-_{(aq)}$$

● 12-5. Use the equation above to determine if a carbonate ion, $CO_3^{2-}{}_{(aq)}$, is an acid or a base. Explain your answer.

12-5. Carbonate ion is a base; it takes a hydrogen ion from a water molecule.

Ammonium chloride, NH_4Cl, is one example of a salt that forms an acidic solution. The chloride ion, $Cl^-_{(aq)}$, neither releases nor takes hydrogen ions. It's the ammonium ion, $NH_4^+{}_{(aq)}$, that releases a hydrogen ion to water molecules to form $H^+_{(aq)}$.

$$NH_4^+{}_{(aq)} \longrightarrow NH_3{}_{(aq)} + H^+_{(aq)}$$

☆ **12-6. How is an acidic solution produced when ammonium chloride, NH_4Cl, dissolves in water?**

12-6. Dissolved ammonium ions, $NH_4^+{}_{(aq)}$, release hydrogen ions.

In the example above, an ammonium ion, $NH_4^+{}_{(aq)}$, can release a hydrogen ion to water molecules. When a water molecule takes a hydrogen ion, it is acting like a base.

Earlier in this activity, you learned that a hydrogen carbonate ion, $HCO_3^-{}_{(aq)}$, can take a hydrogen ion from a water molecule. The water molecule releases a hydrogen ion, thus acting like an acid.

If there is an acid around releasing hydrogen ions, water molecules will tend to take them. If there is a base around taking hydrogen ions, water molecules will tend to release them. So water can act like either an acid or a base.

Even without acids or bases present, water molecules come apart, according to the equation at the top of page 56.

$$H_2O \longrightarrow H^+_{(aq)} + OH^-_{(aq)}$$

Remember that the hydroxide ion is a base. So it will take hydrogen ions and form water.

$$H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O$$

This is the reverse of the other reaction. You can save time and trouble by writing the equation once with reverse arrows.

$$H_2O \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$

Both reactions — forward and reverse — take place zillions of times a second. But only a couple of molecules per billion react at any one instant in pure water. So the reactions don't make water unfit to drink.

12-7. When a water molecule releases hydrogen ions, it acts like an acid; when it takes hydrogen ions, it acts like a base.

☆ **12-7. How can water act like an acid? Like a base?**

ACTIVITY 13: FREEZING AND BOILING

ACTIVITY EMPHASIS: When a solute is dissolved in a solvent, the boiling point of the resulting solution is higher than that of the pure solvent and the freezing point is lower. The size of the elevation or depression depends on the number of solute particles in the solution.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter.

The thermometer should measure to −10° C.

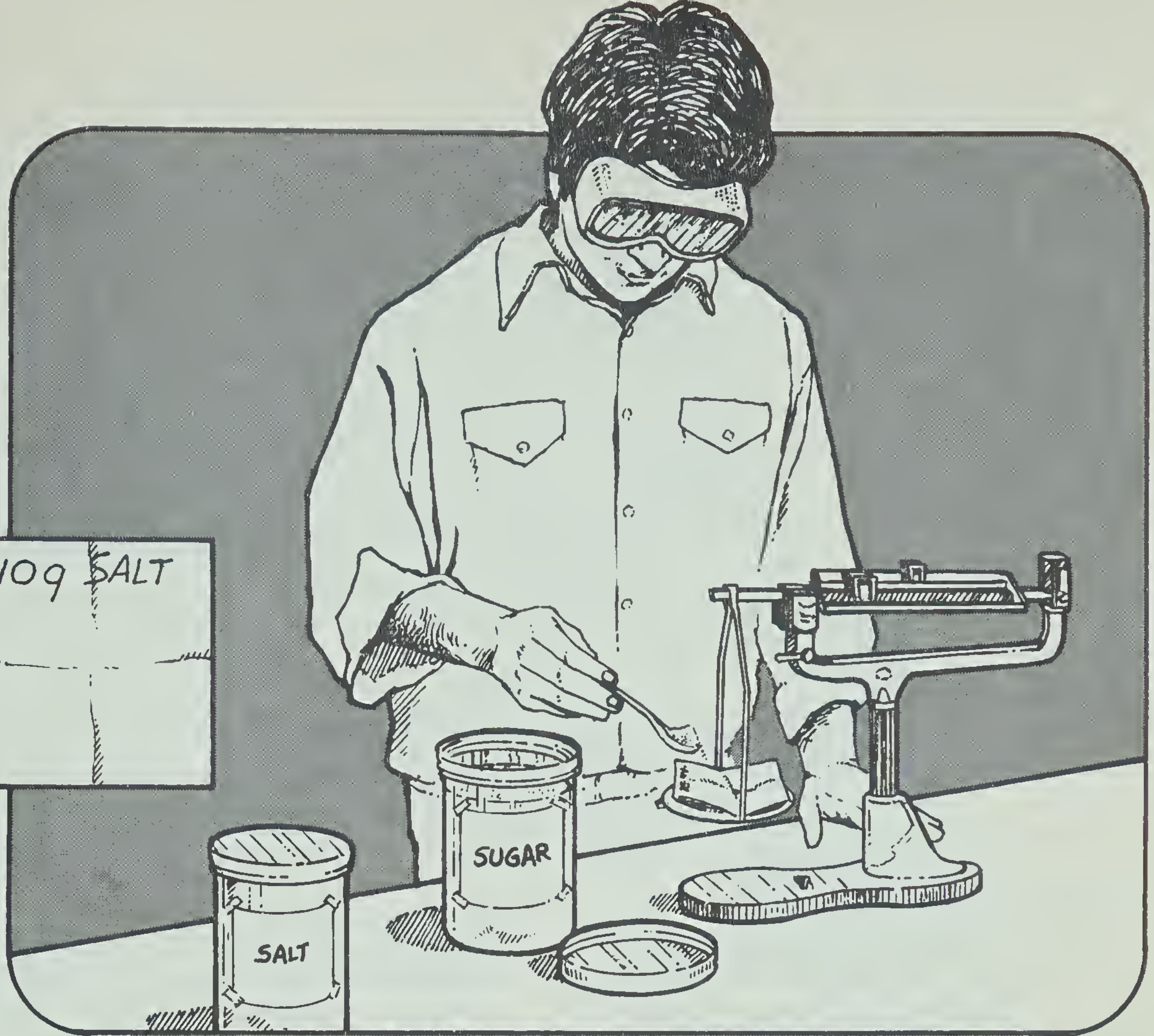
You know that pure water must be at a certain temperature to boil and at a certain temperature to freeze. In this activity you'll see how these "certain temperatures" change when substances are dissolved in the water.

To start, you'll need about thirty minutes and these materials.

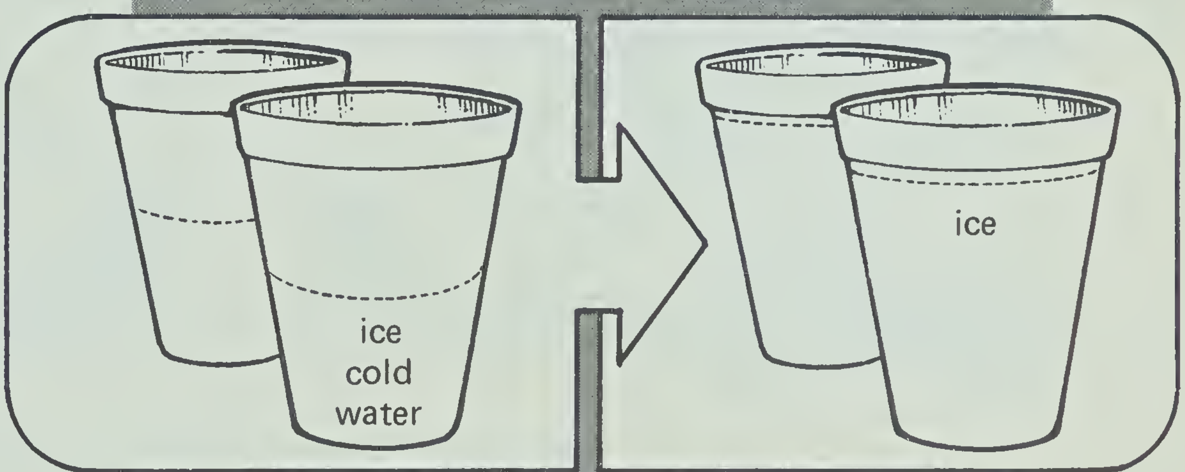
container of crushed ice and ice water
2 Styrofoam cups
2 stirring rods
thermometer
sugar
salt
3 paper squares
grease pencil
balance
safety goggles

A. Label the paper squares as shown.

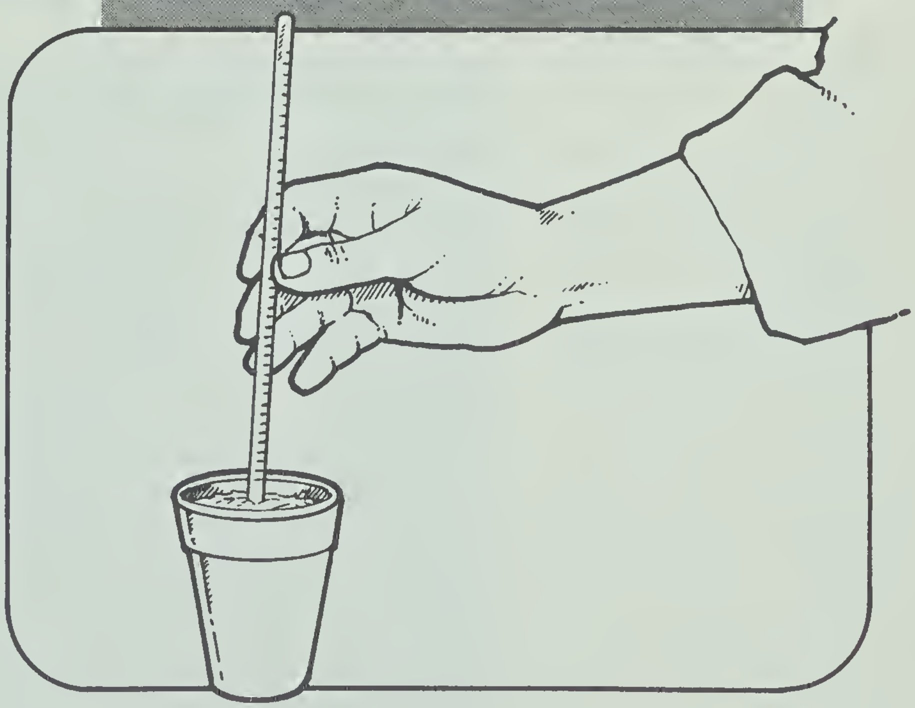
B. On each paper, weigh out that amount of salt or sugar. Set the three samples aside to use later.



C. Half fill each cup with ice cold water from the container. Then fill each cup almost full with ice.

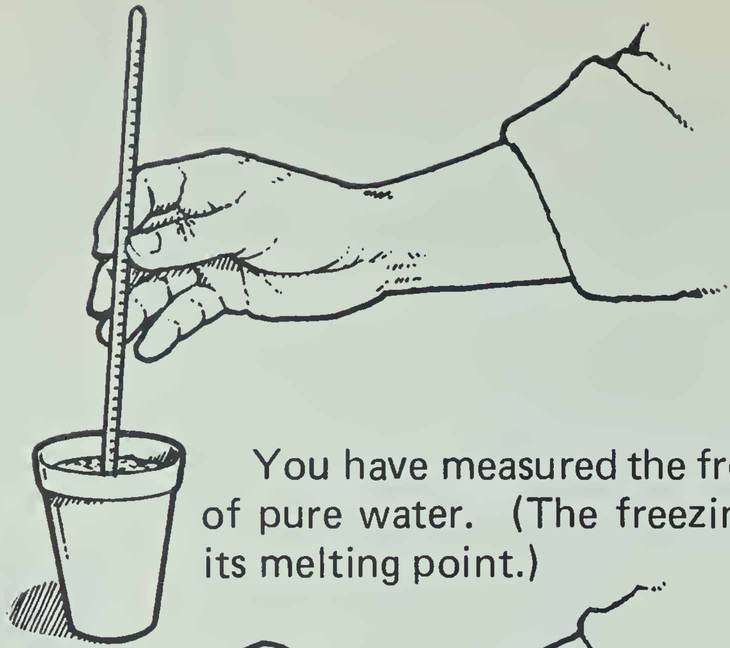


D. Use the thermometer to measure the temperature of the water in each cup.



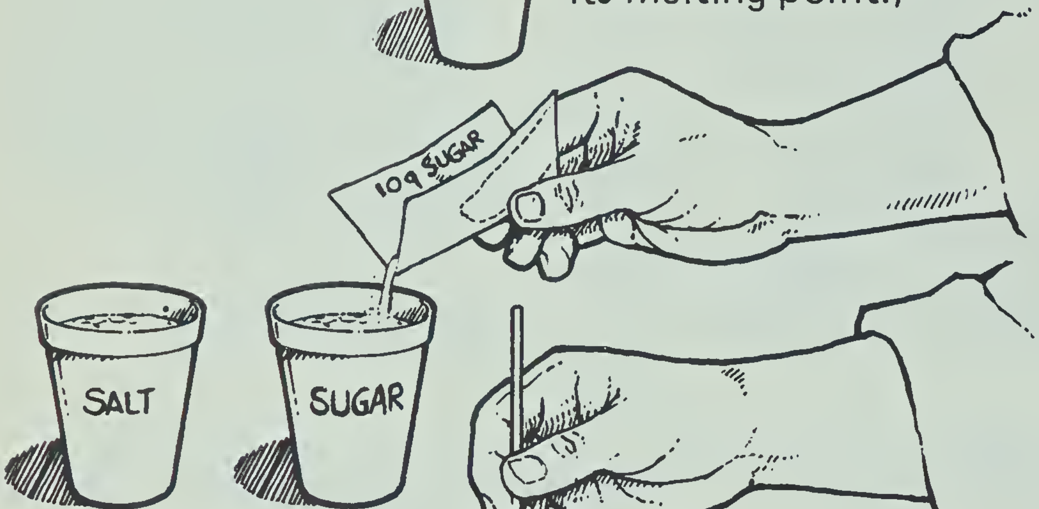
- 13-1. What was the temperature in each cup?

13-1. They should both be the same — about 0° C.



E. If the two temperatures were not the same, measure them again more carefully. When your measurements are the same, continue the investigation.

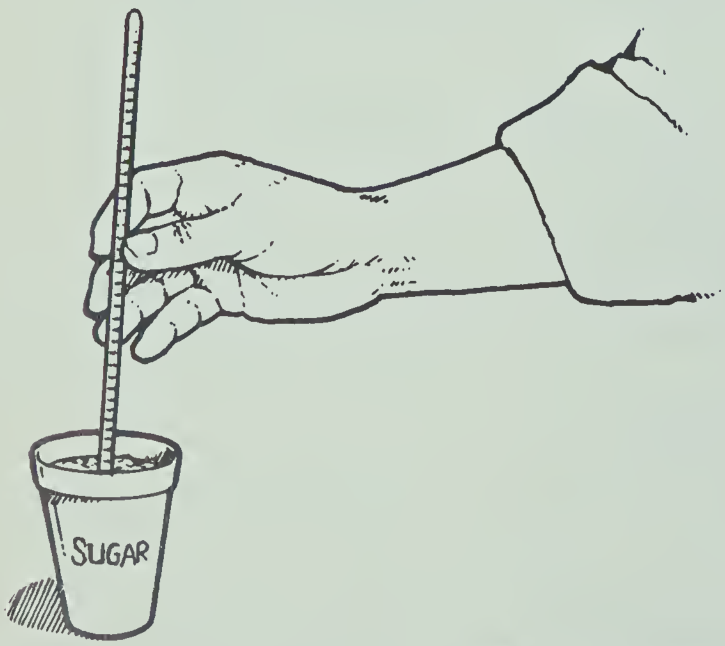
You have measured the freezing temperature, or *freezing point,* of pure water. (The freezing point of a substance is the same as its melting point.)

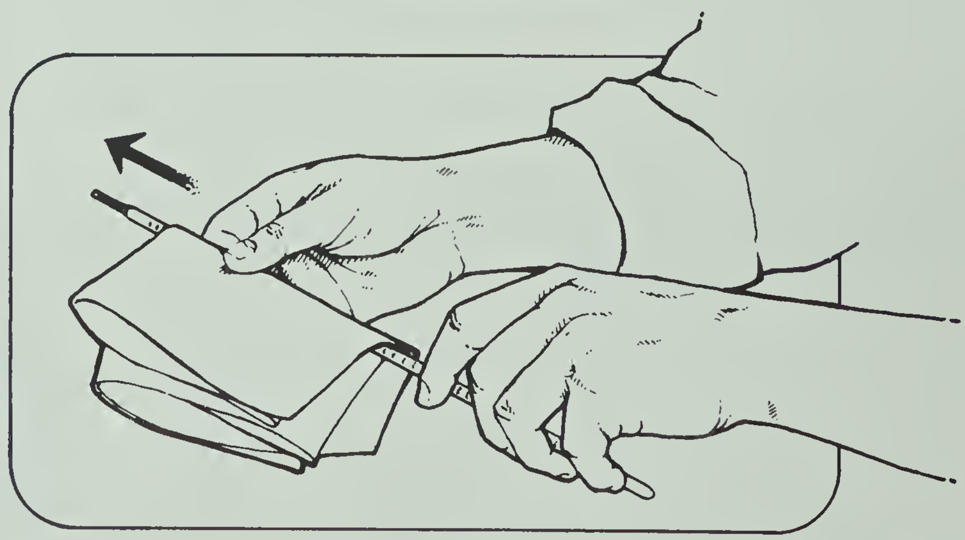


F. Label the cups *salt* and *sugar*. Carefully pour the 10 g of salt into the water in the salt cup. Carefully pour the 10 g of sugar into the water in the sugar cup.

G. With one stirring rod, thoroughly mix the salt and water. With the other stirring rod, thoroughly mix the sugar and water.

H. Now measure the temperature of the solution in each cup. Wipe off the thermometer before you put it into the second cup.





● 13-2. What is the temperature of the sugar solution? The salt solution?

13-2. About –0.5° C; about –5° C

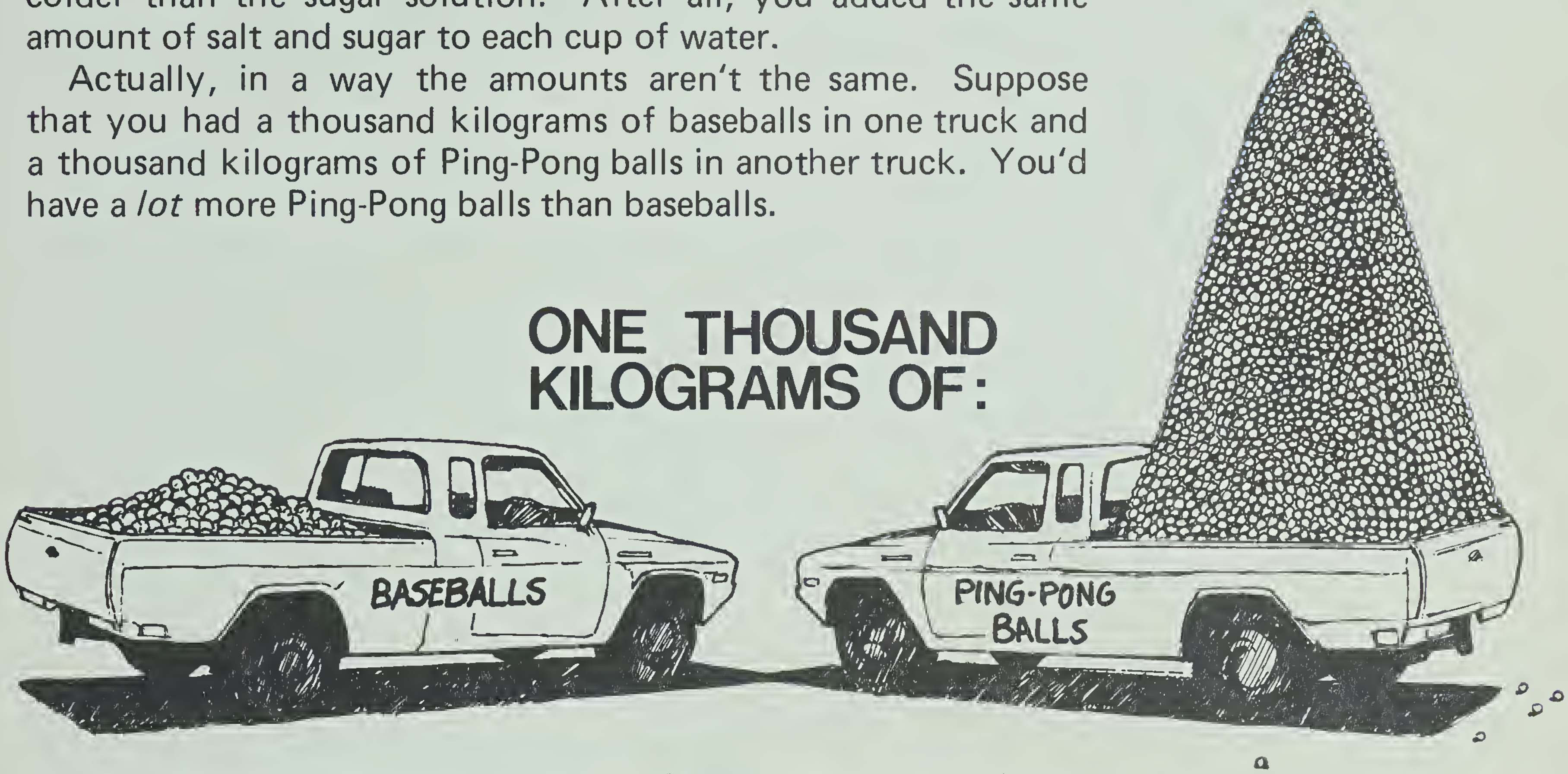
You probably found that the salt solution is a lot colder than the sugar solution. And both are colder than the pure water was. The freezing point of a solution is always lower than the freezing point of a pure liquid.

☆ **13-3. Would the freezing point of an alcohol-in-water solution be higher or lower than the freezing point of pure water?**

13-3. Lower

You're probably wondering why the salt solution is so much colder than the sugar solution. After all, you added the same amount of salt and sugar to each cup of water.

Actually, in a way the amounts aren't the same. Suppose that you had a thousand kilograms of baseballs in one truck and a thousand kilograms of Ping-Pong balls in another truck. You'd have a *lot* more Ping-Pong balls than baseballs.



It's the same with sugar and salt. Sugar molecules are larger and heavier than the ions in salt. Look at Figure 13-1 below.

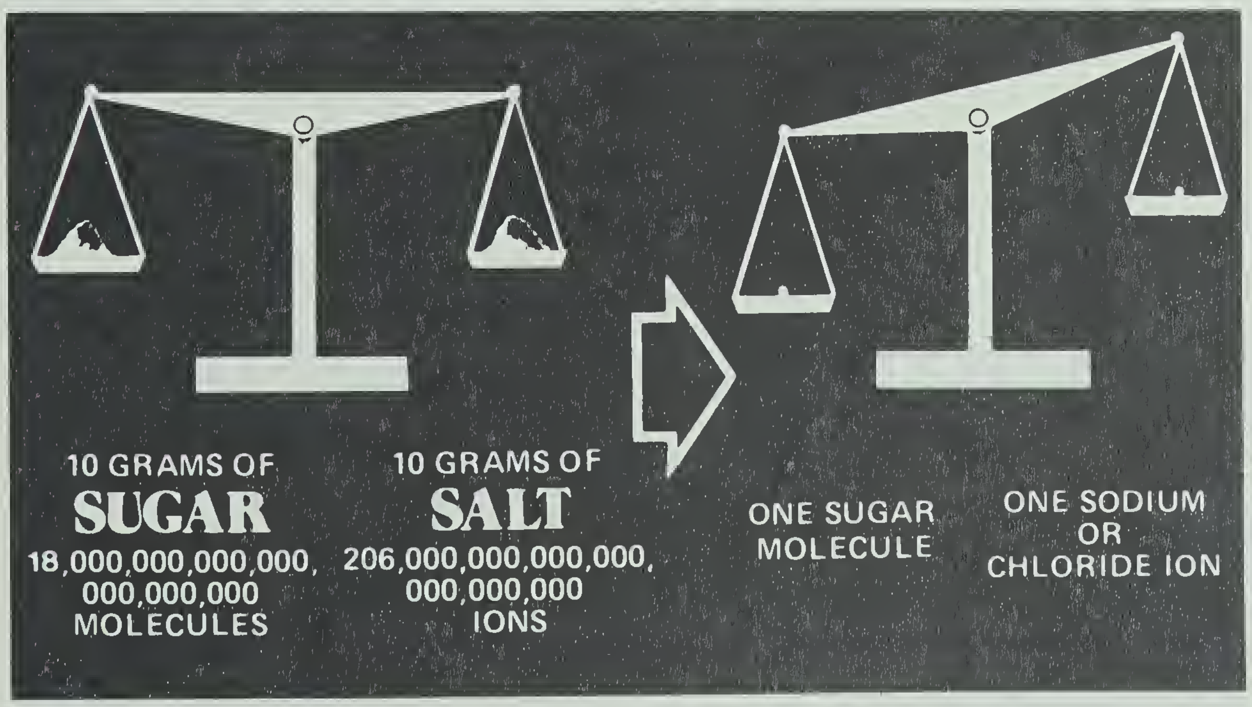


Figure 13-1

13-4. About twelve times

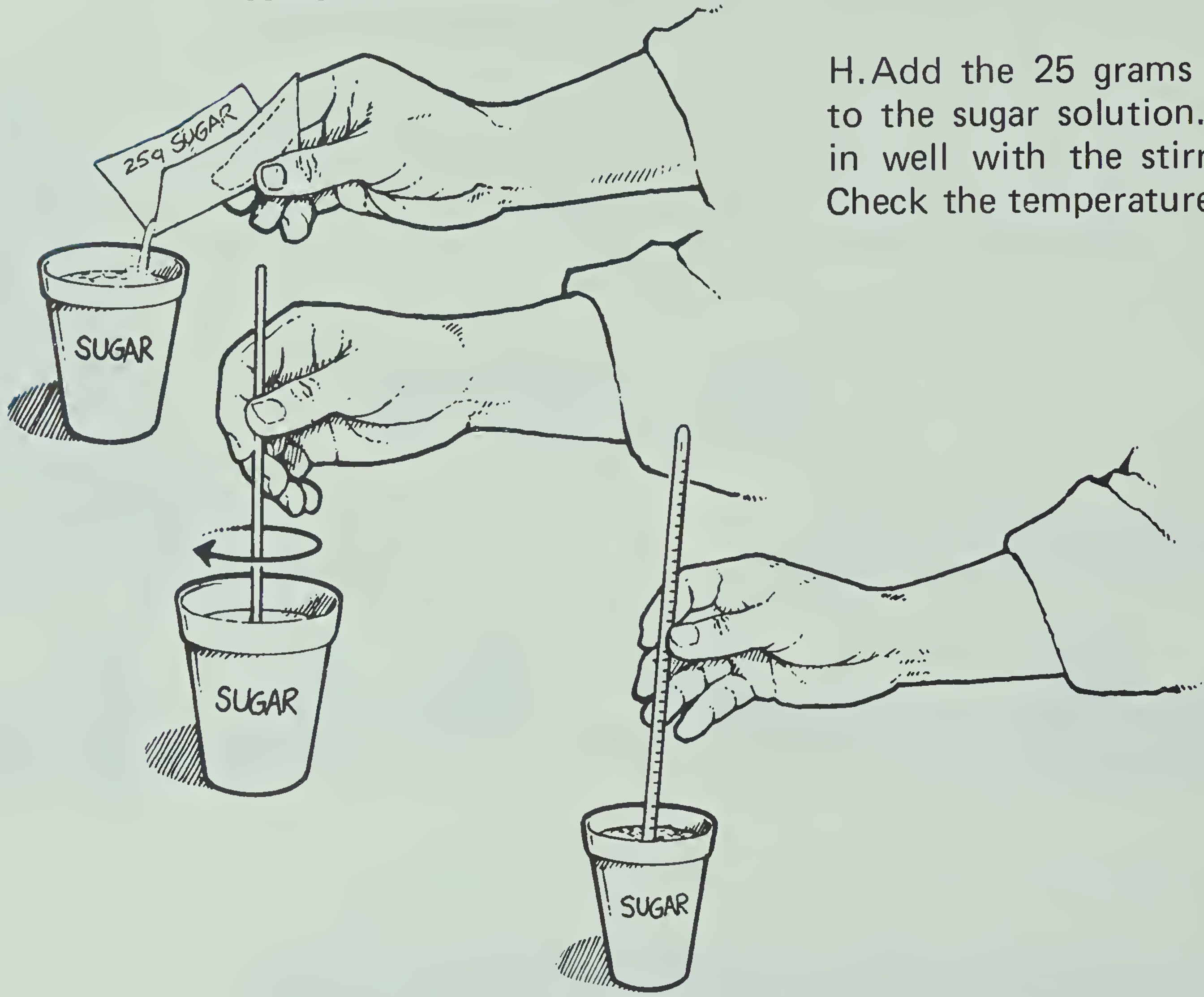
● 13-4. How many times more particles are there in ten grams of salt than in ten grams of sugar?

It's the *number* of particles (molecules or ions) dissolved in the water that determines how much the freezing point is lowered. The more particles, the lower the temperature.

13-5. You could add more sugar to the solution.

● 13-5. How could you lower the freezing point of the sugar solution even more?

H. Add the 25 grams of sugar to the sugar solution. Stir it in well with the stirring rod. Check the temperature now.



13-6. It got lower.

● 13-6. What happened to the temperature after you added the sugar?

It would take about a hundred and twenty grams of sugar to lower the freezing point of water as much as the ten grams of salt did. That's twelve times as much sugar as salt.

13-7. Fifteen grams of Substance A contains more (five times as many) particles than 15 grams of Substance B.

☆ **13-7. If 15 grams of Substance A are added to pure water, the solution freezes at −10°C. When the same amount of Substance B is added to pure water, the solution freezes at −2°C. What does this difference mean in terms of the number of particles in the two substances?**

If you did Activity 7, you may have noticed that the water in the sugar solution boiled far above water's usual boiling temperature (or *boiling point*) of 100°C. The boiling point kept getting higher and higher because more and more water was evaporating. The effect was the same as keeping the amount of water the same and adding more sugar.

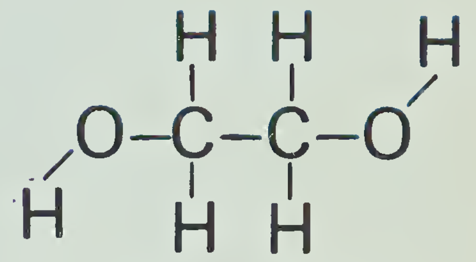
Boiling points are raised, rather than lowered, by particles in solution. The more particles, the more the boiling point is raised.

☆ **13-8. Why do ten grams of salt lower the freezing point of water and raise its boiling point about twelve times more than ten grams of sugar?**

13-8. There are twelve times more particles in the ten grams of salt.

In cold climates, antifreeze is added to the cooling system of car engines. The antifreeze lowers the freezing point of the water in the cooling system. Salt or sugar would also lower the freezing point. But salt would corrode the insides of the engine, and sugar would eventually clog the cooling system.

Instead, antifreeze is made with *ethylene glycol* [ETH-uh-leen GLY-cawl], which won't harm the engine. Ethylene glycol works under the same principle as salt or sugar — the number of particles. Ten grams of ethylene glycol has about half as many particles in solution as ten grams of salt and thus about six times as many as ten grams of sugar.



Ethylene glycol

● 13-9. What happens to the freezing point of the water in the cooling system of a car when ethylene glycol is added to it? What happens to the boiling point?

13-9. The freezing point is lowered; the boiling point is raised.

Since ethylene glycol has six times more particles per gram than sugar does, it will lower the freezing point of water six times more than the same amount of sugar.

● 13-10. How many times more effective is ethylene glycol at raising the boiling point of water than sugar is?

13-10. Six times more effective

☆ **13-11. Five grams of a green solid, added to 100 ml of water, raises the boiling point to 102.5°C. The same amount of a white solid, added to 100 ml of water, raises the boiling point to 100.5°C. Explain this difference in terms of the number of particles in the green and white solids.**

13-11. The green solid contains more (five times as many) particles than the white solid.

EXCURSION

ACTIVITY 14: PLANNING

Activity 15 **Page 63**

Waxing and Polishing

The best polishing creams are suspensions of a waxy substance in water. And they're easy to make. This excursion gives you a secret recipe invented many years ago. Each secret ingredient and its purpose will be revealed. The polish will give your shoes, belt, or handbag a better gloss than any wax polish you've ever used!

Activity 16 **Page 69**

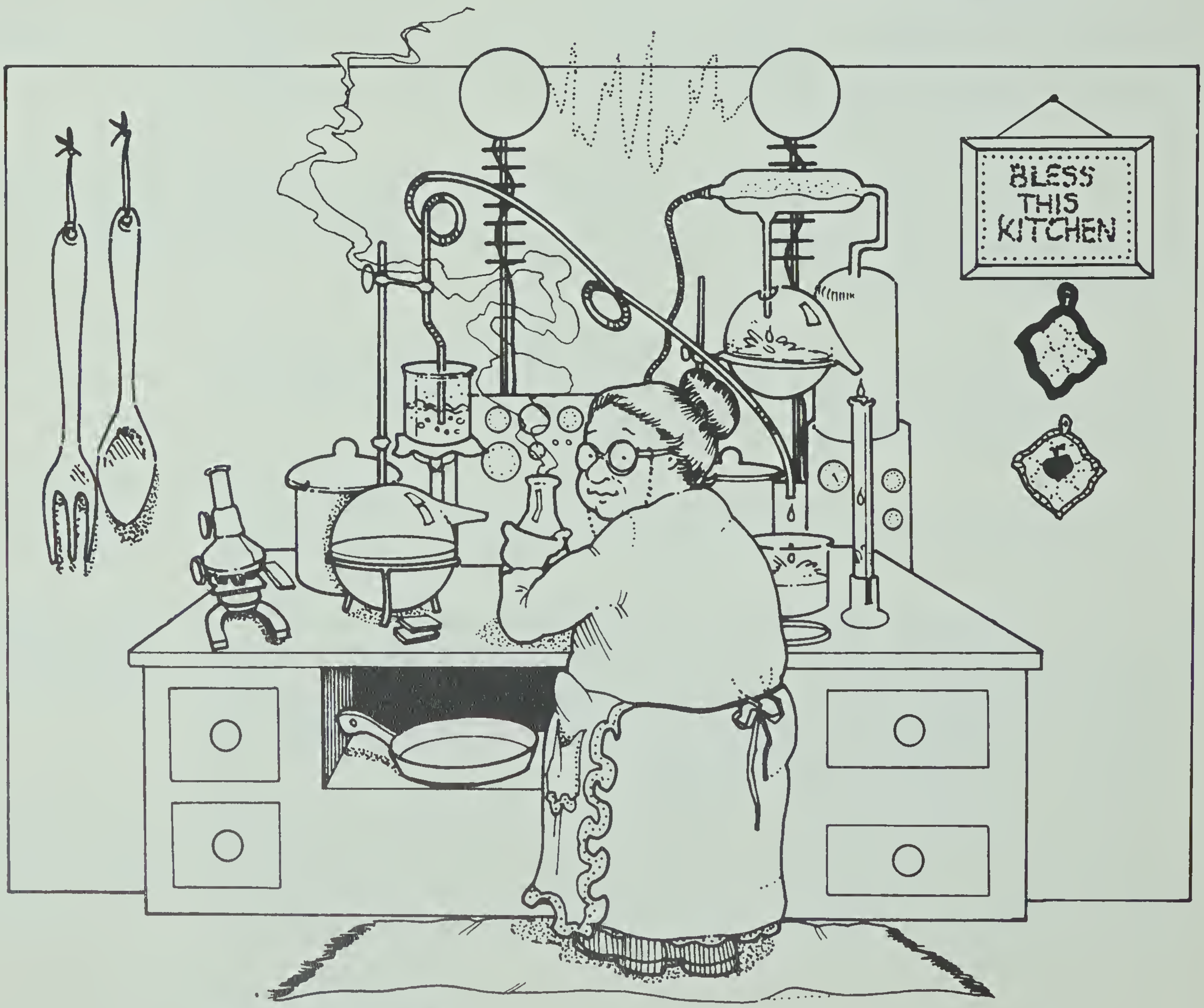
More on Suspensions

You've probably heard that oil and water don't mix. But if you do it right, you can make them into an emulsion. In this excursion, you'll make a foam and some gelatin. You'll also get a recipe for homemade mayonnaise and learn about butter.

Activity 17 **Page 74**

A Homemade Indicator

You can do more with cabbage than just eat it. In this excursion, you'll use cabbage to make your own indicator. Then you'll find out what happens when you test various acidic and basic solutions.



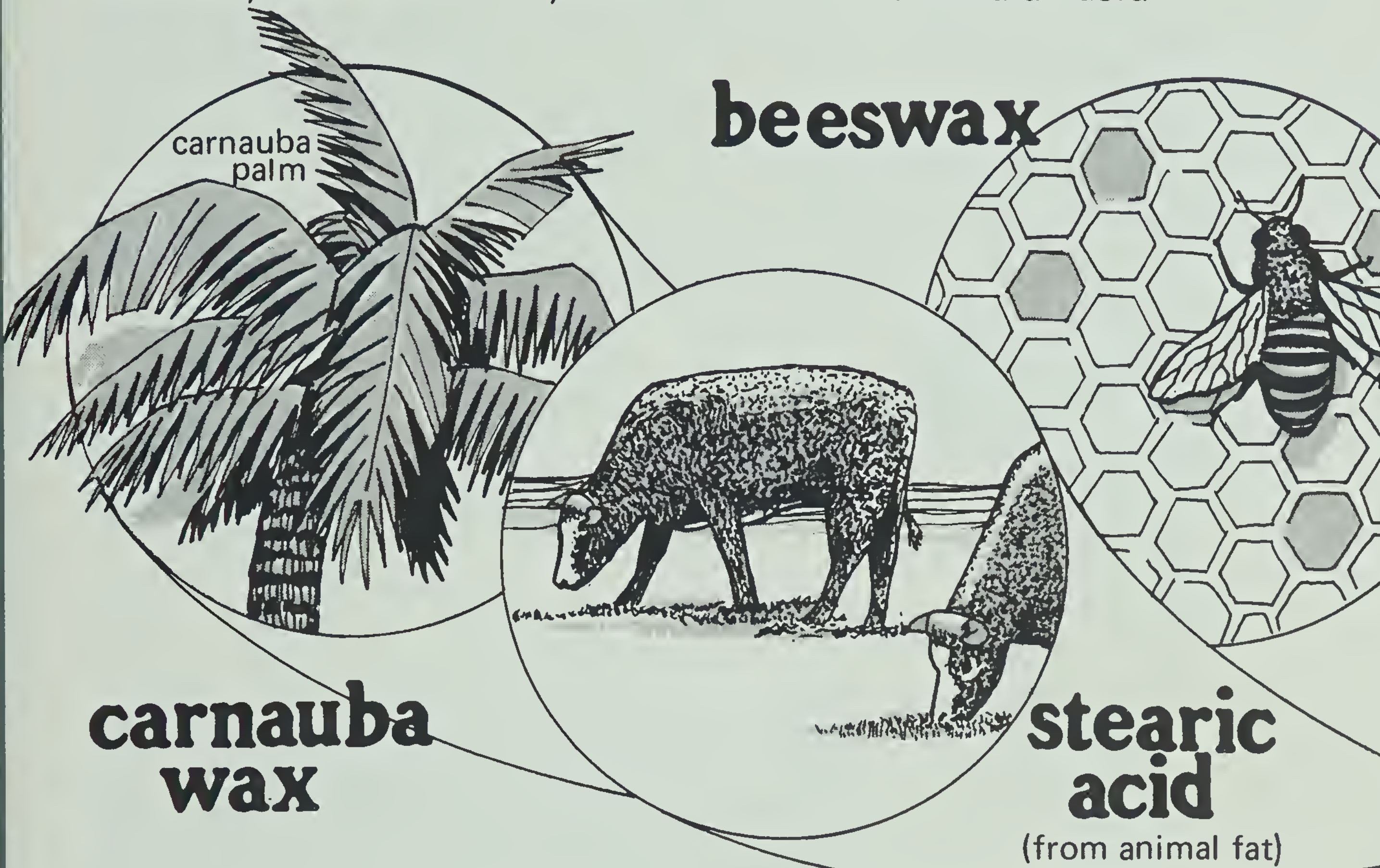
ACTIVITY 15: WAXING AND POLISHING

ACTIVITY EMPHASIS: A suspension of stearic acid particles (a wax) can be prepared and used as a polish.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

When you wax a piece of furniture or your shoes or anything else, you put a thin layer of wax on the surface. When the layer of wax is polished, it reflects light better and makes the object more attractive.

Since waxes are somewhat hard, the problem is how to get a thin layer on an object with the least trouble. Suppose you wanted to shine your shoes. You might not think of using candle wax. But good candles are made from carnauba wax, beeswax, and stearic acid, which is both a wax and an acid.



Carnauba wax comes from the leaves of the carnauba palm tree from Brazil. It's the best wax for shoes and furniture because it's hard and spreads easily. But it's expensive.

Beeswax is produced naturally by bees. It's not very good for polishing surfaces because it's too soft.

Stearic acid is a waxy solid produced from animal fat. It's good for polishing and is not nearly as expensive as carnauba wax.

● 15-1. What properties of a wax make it useful as a polish?

15-1. Hardness and spreadability

You'll use stearic acid to make your polish. This substance won't harm leather, wood, or plastic items in any way. The idea is to get the wax into very, very small particles, spread the particles on your shoes or handbag or table, and then shine it. You'll need to make a permanent suspension of stearic acid particles in water.

15-2. Very, very small

• 15-2. What size will the stearic acid particles need to be to make the suspension permanent?

To get the stearic acid particles small enough to be in permanent suspension, you will need some ammonium stearate. You will make this from ammonia water and stearic acid. Also, to help the shine a bit, you will add a small amount of shellac. The polish will take about fifty minutes to make. You'll need a partner and these materials.

Use shellac normally sold in stores as a solution in alcohol. The best results come from 3-pound cut, but 4- or 5-pound cut will work.

Use 15M ammonia solution, which is 28% NH_3. Caution students not to breathe the fumes.
Alcohol can be used to clean up test tubes and jars.

2 small test tubes with stoppers
grease pencil
metric ruler
dropper
wooden tongue depressor
shellac
ammonia water, 15M, in dropping bottle
15 g stearic acid
3 250-ml beakers
2 Bunsen burners or other heat sources
2 ring stands, rings, and wire gauzes
test-tube rack
2 soft cloths
small jar with lid
alcohol for cleanup
safety goggles

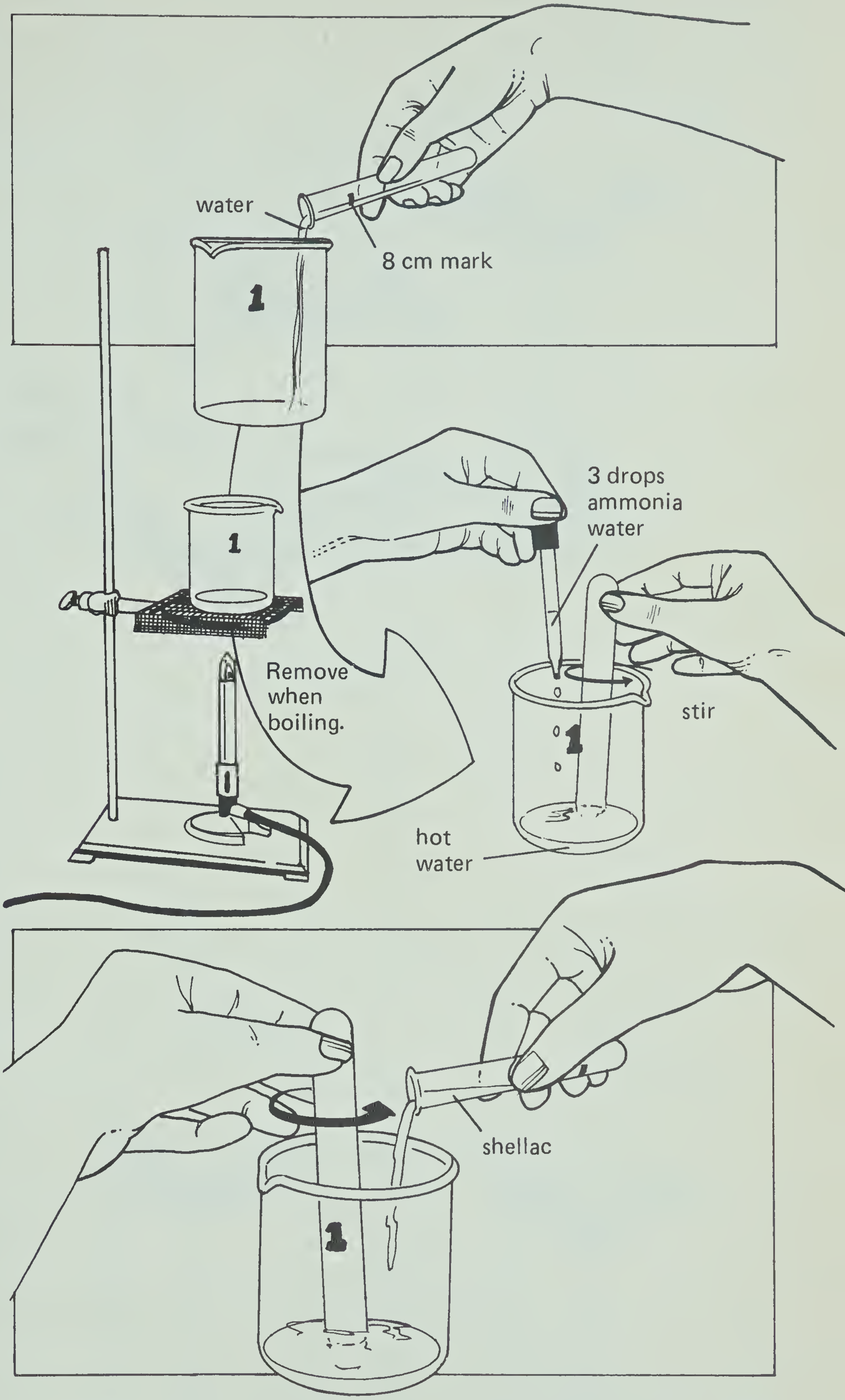
CAUTION	This procedure is fairly complicated. Before you begin Step A, read through *all* the steps. Decide what each partner will do. Wear your safety goggles.

A. With the grease pencil, mark one test tube 3 cm from the bottom. Fill it to the mark with shellac. Stopper the test tube so that the shellac does not harden.

B. Mark the other test tube 8 cm from the bottom, and fill it to the mark with water. Label one 250-ml beaker with a *1*. Pour the water from this test tube into Beaker 1.

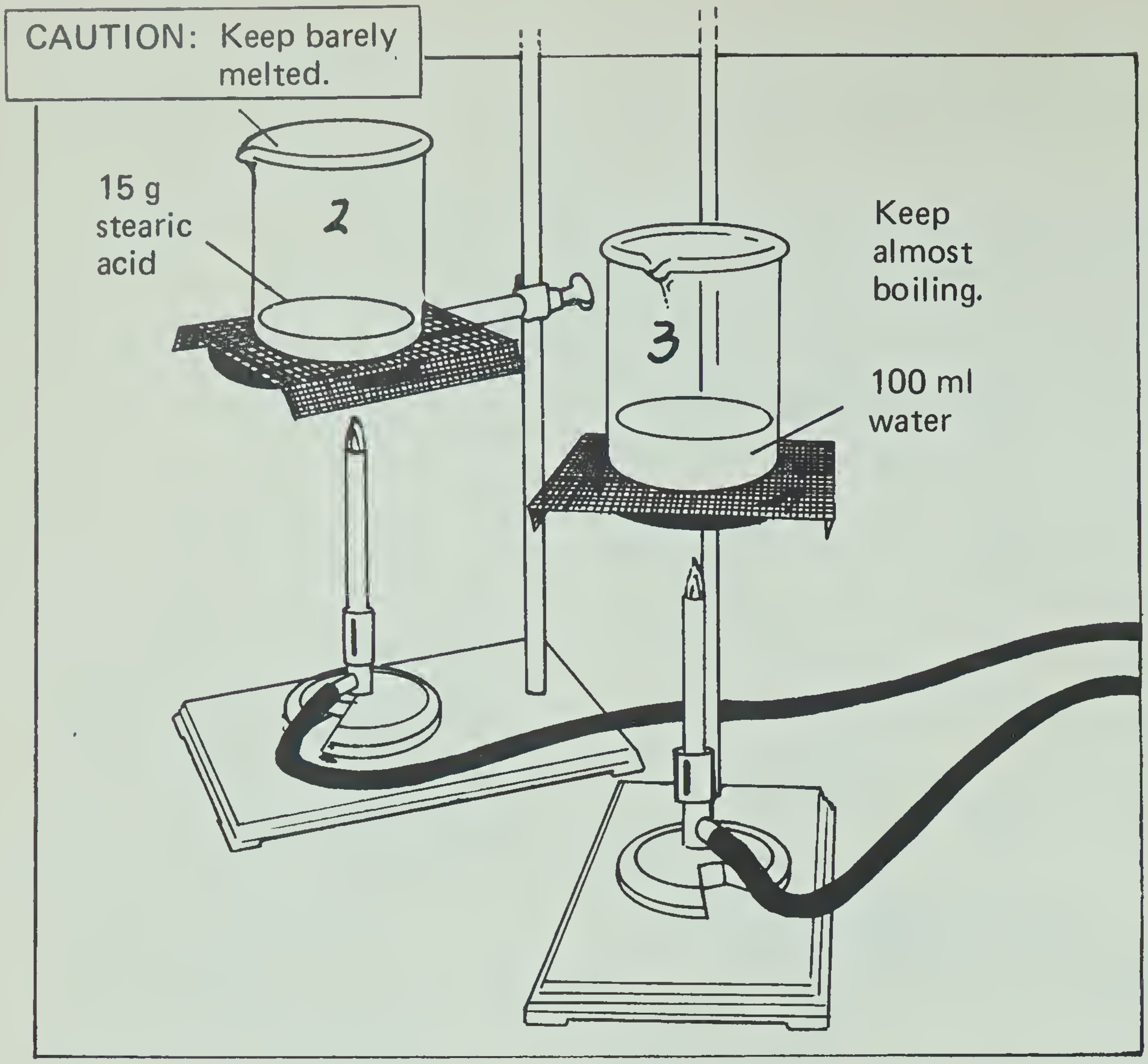
C. Light one burner and heat the water in Beaker 1 to boiling. Then remove the beaker from the heat, add 3 drops of ammonia water, and stir the mixture with the tongue depressor.

D. Remove the stopper from the test tube, and slowly add all the shellac to Beaker 1, stirring the mixture well as you do so. Stopper the shellac test tube again so that you can clean it easily later. Set Beaker 1 aside to cool.

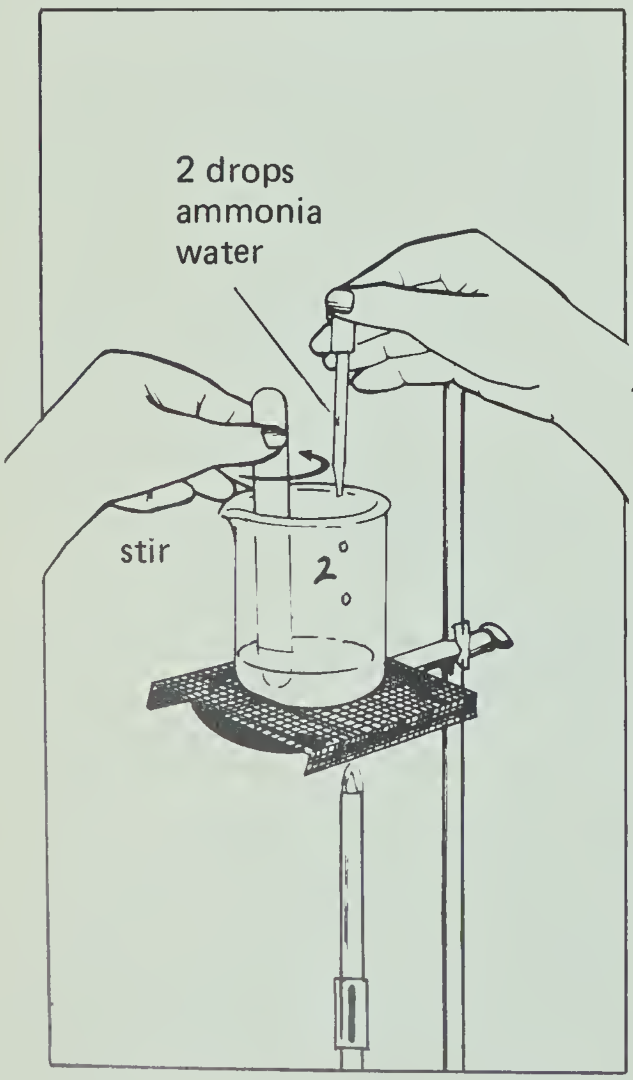


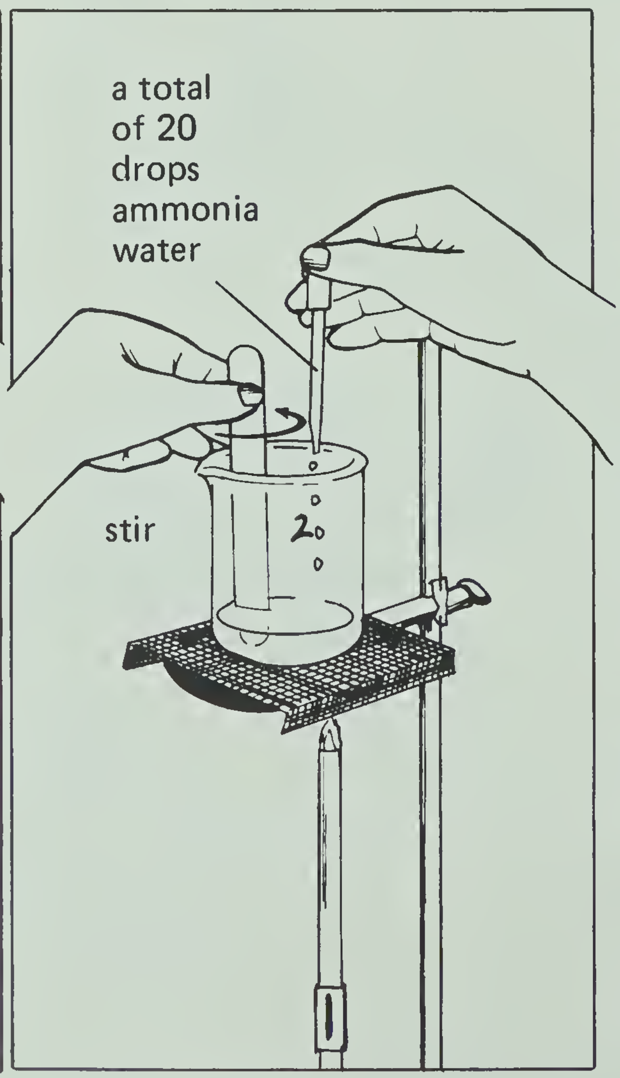
15-3. Shellac, ammonia water, and water

- 15-3. What ingredients are now in Beaker 1?



E. Get the other two 250-ml beakers. Label one *2* and one *3*. Put 15 g of stearic acid into Beaker 2. Put 100 ml of water into Beaker 3. Light both burners, and heat both beakers. Use a low flame on the stearic acid if you're using a Bunsen burner. Heat the stearic acid in Beaker 2 gently until it just melts. *Don't* let it boil. The water needs to be very hot – almost boiling.

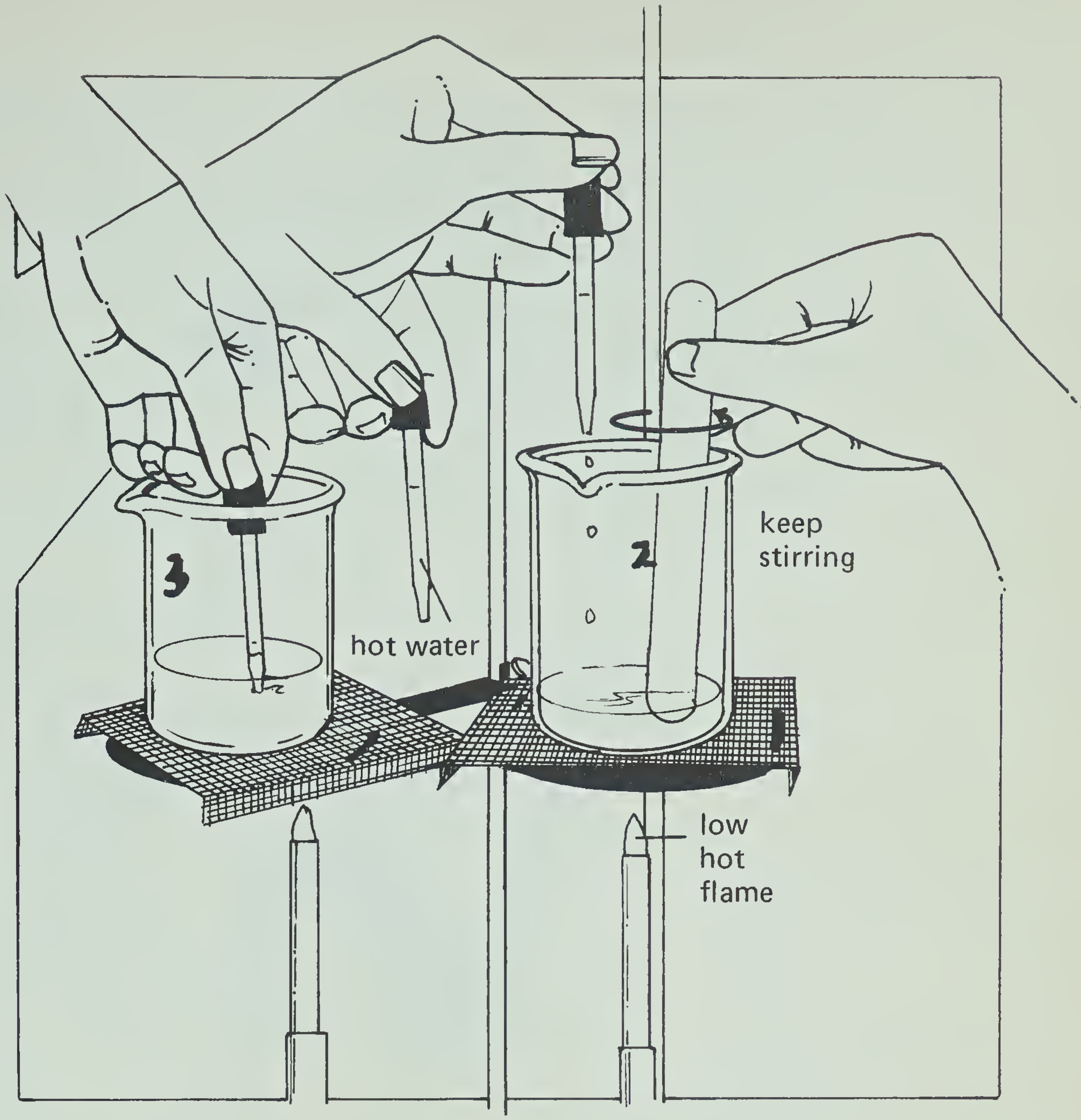




F. After the stearic acid in Beaker 2 has all melted, stir it gently with the tongue depressor and add two drops of ammonia water from the dropping bottle. Keep on stirring, and keep it melted.

G. After that solid – which was ammonium stearate – has dissolved, add two more drops of ammonia water to make more ammonium stearate. Keep on stirring. When that ammonium stearate has dissolved, add two more drops of ammonia water and continue to stir. Keep doing that until you have added 20 drops of the ammonia water to Beaker 2. Keep on stirring, and keep it melted.

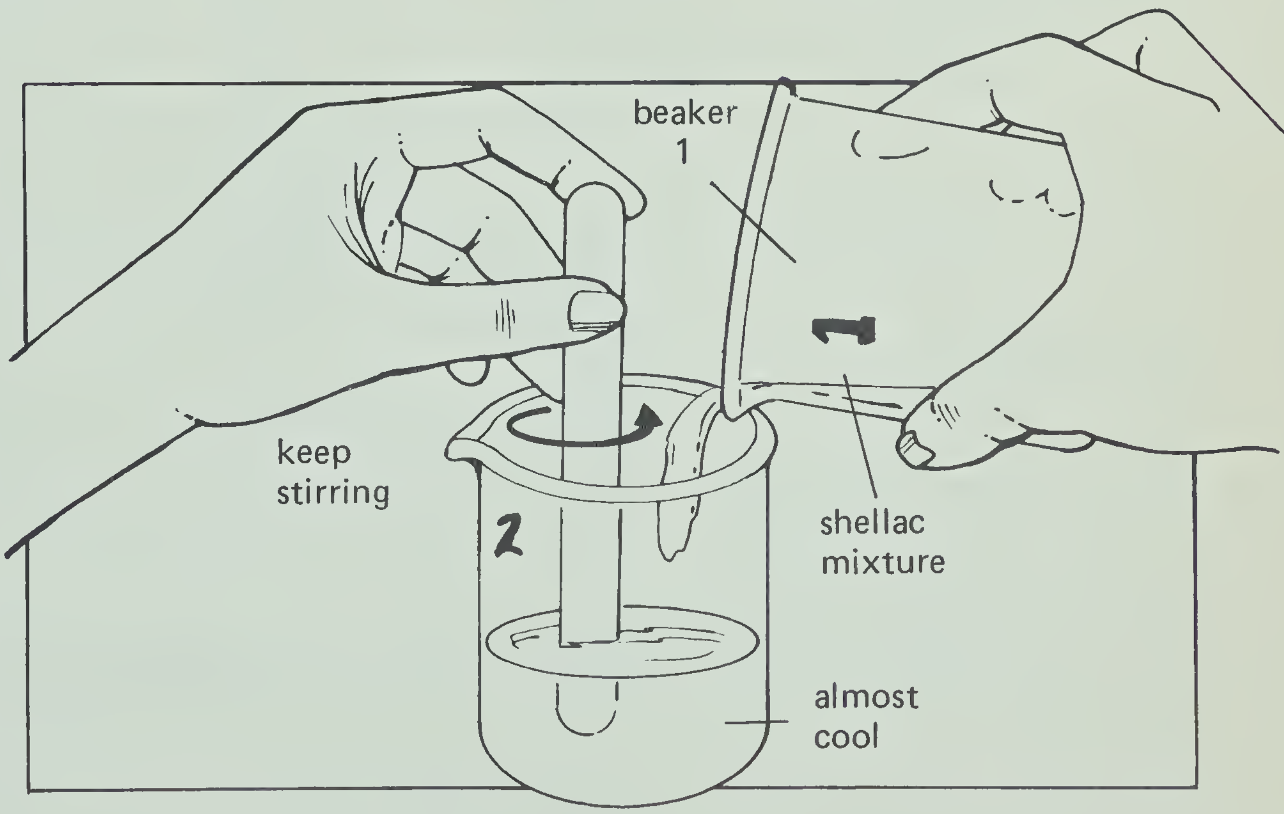
H. Keep on stirring the melted stearic acid and ammonium stearate in Beaker 2. With the dropper, add a few drops of hot, almost boiling, water from Beaker 3. Continue to stir the mixture in Beaker 2.



I. Keep on stirring, and add more water from Beaker 3 little by little, a few drops each time. Keep on stirring until all the water has been added. Do not add the water rapidly — only a few drops at a time. Don't stop stirring.

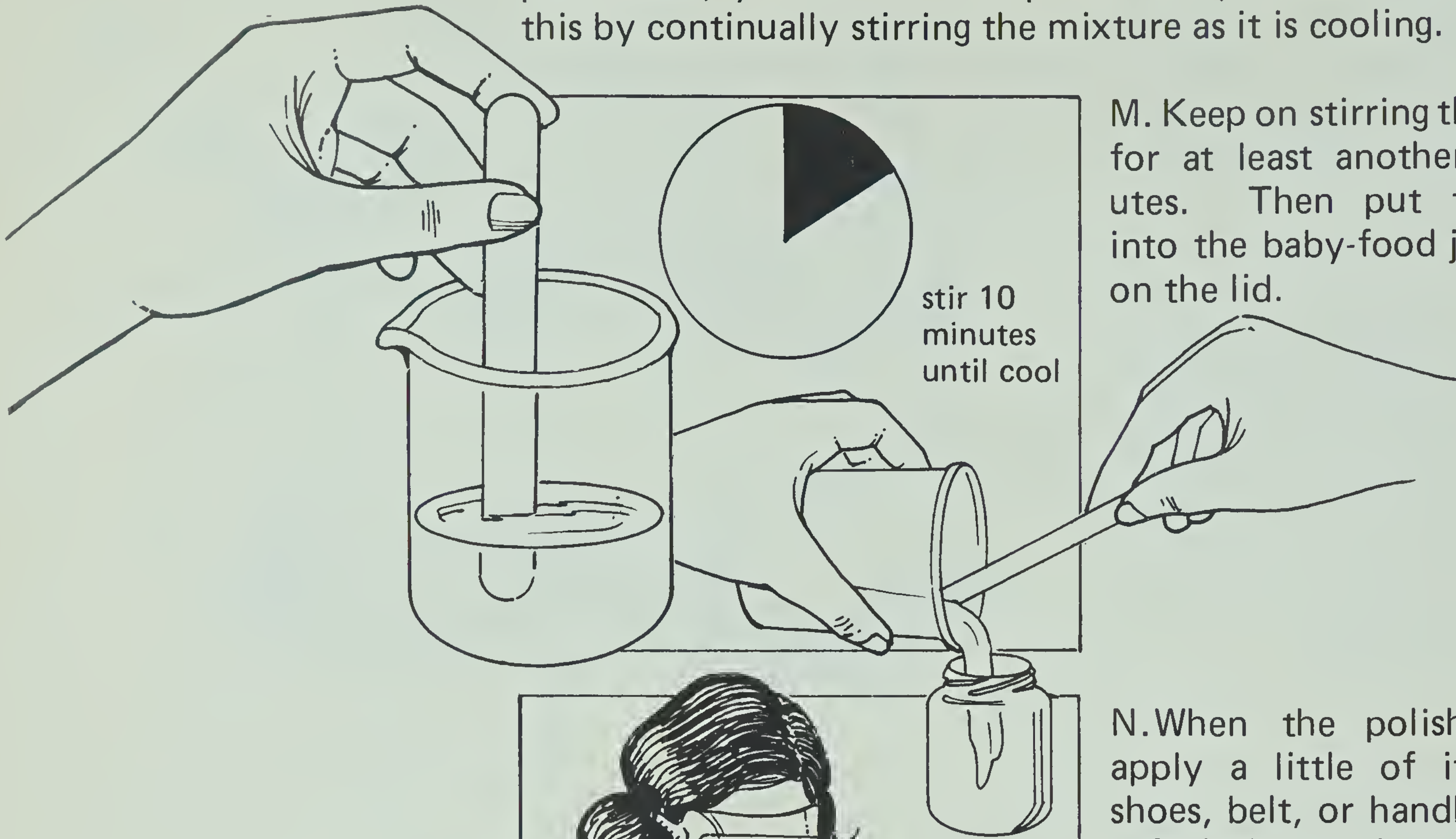
J. If the water drops "pop" when you add them to Beaker 2, it is too hot. Take away the burner for a couple of minutes.

K. By the time all the water has been added, you will have a white, cloudy mixture. Turn off both burners, and keep on stirring the cloudy mixture.



L. Little by little, as you keep stirring, the cloudy mixture will cool and get thicker. When it is almost cool, it will be quite thick. Then stir the shellac—water—ammonia water mixture from Beaker 1 into Beaker 2. Stir it in well.

The material you now have is a polish made up of small particles of stearic acid wax in suspension. To keep this suspension permanent, you need to keep the wax particles small. You do this by continually stirring the mixture as it is cooling.



M. Keep on stirring the mixture for at least another ten minutes. Then put the polish into the baby-food jar and put on the lid.

N. When the polish is cool, apply a little of it to your shoes, belt, or handbag. Use a soft cloth, and choose a smooth section of the leather or plastic. Spread it around well, and let it dry for about five minutes. Then shine the leather or plastic with a clean soft cloth. (The polish may leave white streaks in the wrinkles of older leather or plastic. If this happens, the streaks can be removed by vigorous polishing with a brush.)

O. Use alcohol to clean the shellac and polish from the test tube and beakers.

● 15-4. List the ingredients you mixed to make the polish.

☆ **15-5. What is the purpose of each ingredient in the polish? (Hint: Look at pages 63 through 67 again.)**

● 15-6. What did you do to keep the wax particles in suspension?

15-4. Stearic acid, ammonia water, water, and shellac

15-5. Stearic acid was the wax. Stearic acid and ammonia water reacted to produce ammonium stearate, which made the wax particles small enough to be in permanent suspension. The wax particles were suspended in the water. Shellac was used to help the shine.

15-6. Stirred the mixture for a long time

ACTIVITY 16: MORE ON SUSPENSIONS

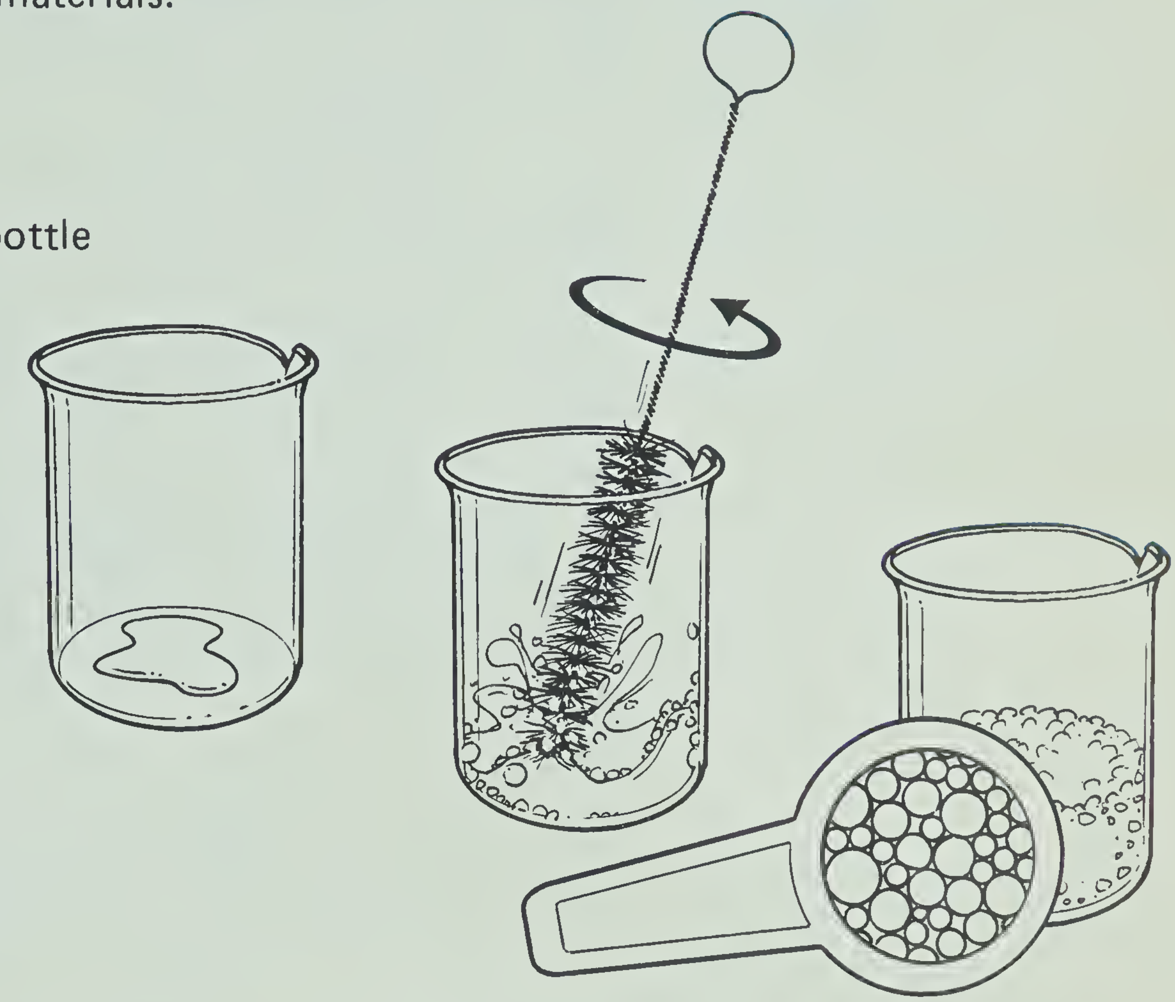
ACTIVITY EMPHASIS: Students examine examples of four kinds of suspensions: liquid in liquid (emulsion), gas in liquid (foam), solid in liquid, and liquid in solid (sol or gel).

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter.

From your work in core, you know that solid dust particles are suspended in the air. Also, tiny bubbles of gas are suspended in such things as meringue, popcorn, and bread. Many of the common suspensions found in the kitchen are made of substances suspended in a liquid, such as water.

You can make a gas-in-liquid suspension, called *foam*, in about five minutes. You'll need these materials.

test-tube brush
250-ml beaker
water in dropping bottle
liquid detergent in dropping bottle
hand lens



A. Put a small amount of soap or detergent into the beaker. Add a little water. The actual amounts of each aren't too important, but don't use much. You can always add more later, if necessary.

B. Whip the mixture with the brush until a lather forms.

C. Use the lens to examine the lather closely.

- 16-1. Describe the lather that you made.

16-1. There are many tiny bubbles all stuck together.

The tiny bubbles of air are suspended in the soapy liquid, making a foam. Shaving lather from a dispenser, whipped cream, and beaten egg white are other examples of gas-in-liquid suspensions.

Another kind of suspension is a liquid-in-liquid suspension. To make one, you'll need about twenty minutes and the following materials.

medium test tube with stopper
cooking oil
liquid detergent in dropping bottle
test-tube rack
10-ml graduated cylinder
safety goggles

A. Put 10 ml of water into the test tube. Then add 5 ml of cooking oil.

B. Stopper the test tube, and shake it several times as hard as you can. Then put it into the test-tube rack.

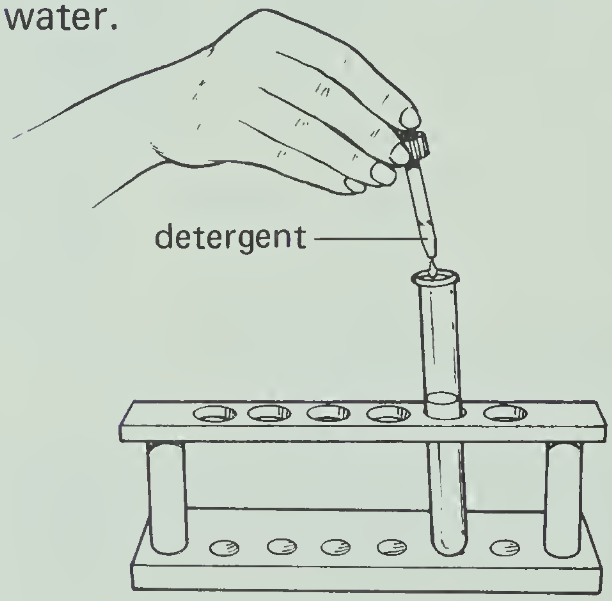
16-2. There are small drops of oil mixed in with the water.

● 16-2. Describe the oil and water mixture just as you stopped shaking it.

16-3. The oil and water have separated.

● 16-3. Describe the oil and water mixture about one minute later.

As soon as you stopped shaking the test tube, you had a temporary suspension of oil droplets in water. But most of the droplets were so big that they couldn't stay in suspension. They joined together in about a minute and rose to the top of the water.



C. Remove the stopper, and add a few drops of detergent. Put the stopper back on. Shake the test tube several times, as hard as you can. Then put it back into the test-tube rack.

16-4. The oil, water, and detergent stayed mixed together.

● 16-4. Describe what happened that time.

If you shook the test tube hard enough and long enough, you made a permanent suspension. A permanent suspension of one liquid in another, such as oil in water, is called an *emulsion*. But detergent was needed to make your emulsion. The detergent is called an *emulsifying agent*.

When you wash clothes or dishes, the detergent emulsifies the oil and grease, making a permanent suspension of them in the wash water.

You wouldn't want to eat detergent, but you can eat some emulsifying agents. Mayonnaise, another liquid-in-liquid suspension, is an emulsion of vegetable oil in vinegar. Egg yolk is the emulsifying agent. Many salad dressings that don't separate are also emulsions.

Here's a recipe for making your own mayonnaise at home.

HOMEMADE MAYONNAISE

2 egg yolks	250 ml salad oil
45 ml vinegar	small bowl
2 ml salt	egg beater or electric mixer
1 ml prepared mustard	

Have all ingredients at room temperature. Put 15 ml of the vinegar into the bowl. Add the egg yolks, salt, and mustard. Beat the egg-yolk mixture until it is sticky. Continue beating it, adding the salad oil drop by drop. The emulsion has been formed when the mixture gets thick — usually after about 100 ml of oil have been added. Add the remaining oil more quickly until it has all been added. Then beat in the remaining 30 ml of vinegar.

Your mayonnaise should be light yellow in color, thick, and shiny looking. Cover it, and store it in a refrigerator so it won't spoil too fast.

☆ **16-5. Tell whether each of the following is a gas-in-liquid suspension or liquid-in-liquid suspension. (Check your answers in the text on page 69 and above.)**

A. Mayonnaise

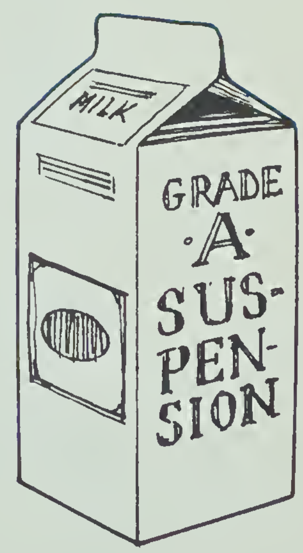
B. Whipped cream

C. Shaving lather

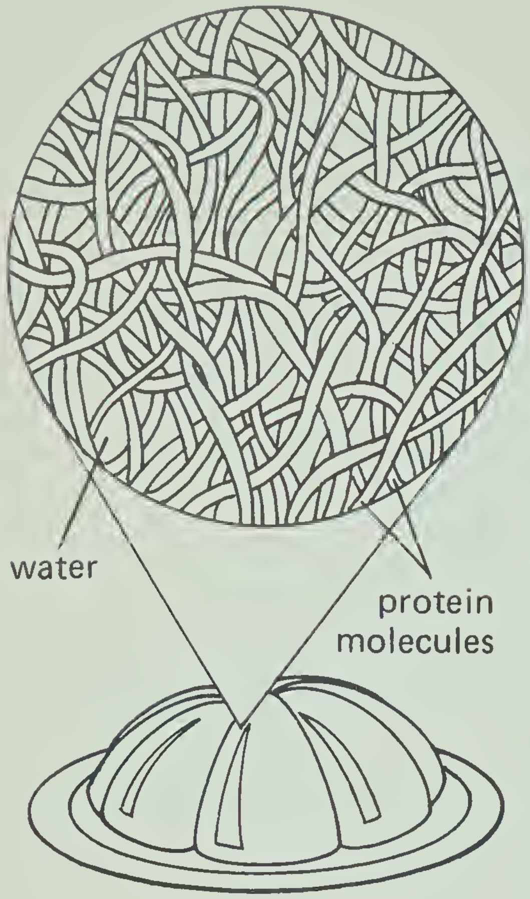
D. French salad dressing

16-5. A. Liquid-in-liquid suspension; B. gas-in-liquid suspension; C. gas-in-liquid suspension; D. liquid-in-liquid suspension

Still another kind of permanent suspension is made of solids suspended in liquids. Milk is a permanent suspension of solid particles, mostly protein particles and butterfat, in a watery liquid. Self-polishing floor wax is a permanent suspension of tiny solid wax particles in a watery liquid. Coffee is a permanent suspension of solid particles in a liquid. But coffee is partly a temporary suspension, since the bottom of the coffee cup sometimes has some solid particles remaining.



There are other examples of solid-in-liquid suspensions. One of the most interesting is gelatin. The solid part in gelatin is made up of long, narrow protein molecules, something like spaghetti, but much smaller.



To get an idea of this, think of a bowl of spaghetti with tomato sauce. If the spaghetti is all tangled up, there will still be spaces all over the place filled with tomato sauce.

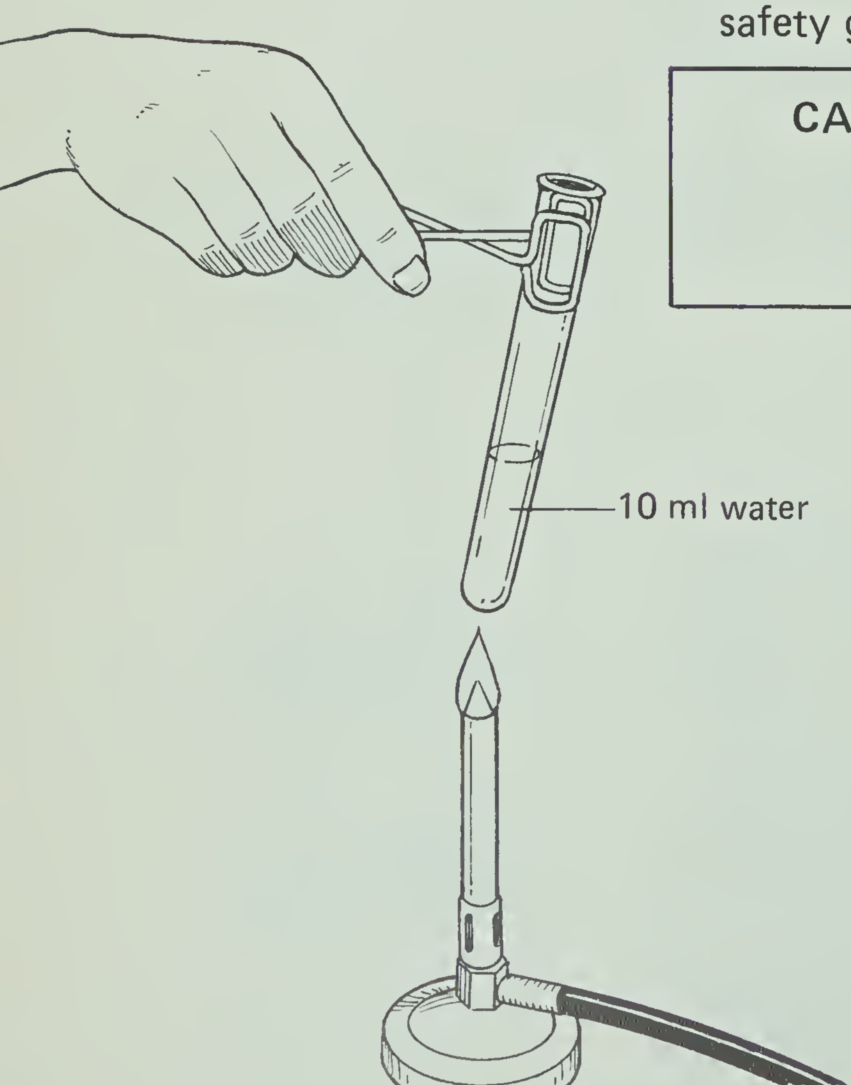
It's the same with gelatin. The solid particles are long protein molecules all tangled up. There is water in the spaces between and among the tangled molecules.

When you make gelatin dessert, you start with small granules of gelatin. These granules are the tangled molecules. When you add water to the granules, they swell as the water works its way between the tangles, pushing them apart. The molecules may get partly untangled too. If you add too much water, they will get untangled and you'll have soup instead of gelatin dessert.

To make your own gelatin, you'll need about fifteen minutes and the following materials.

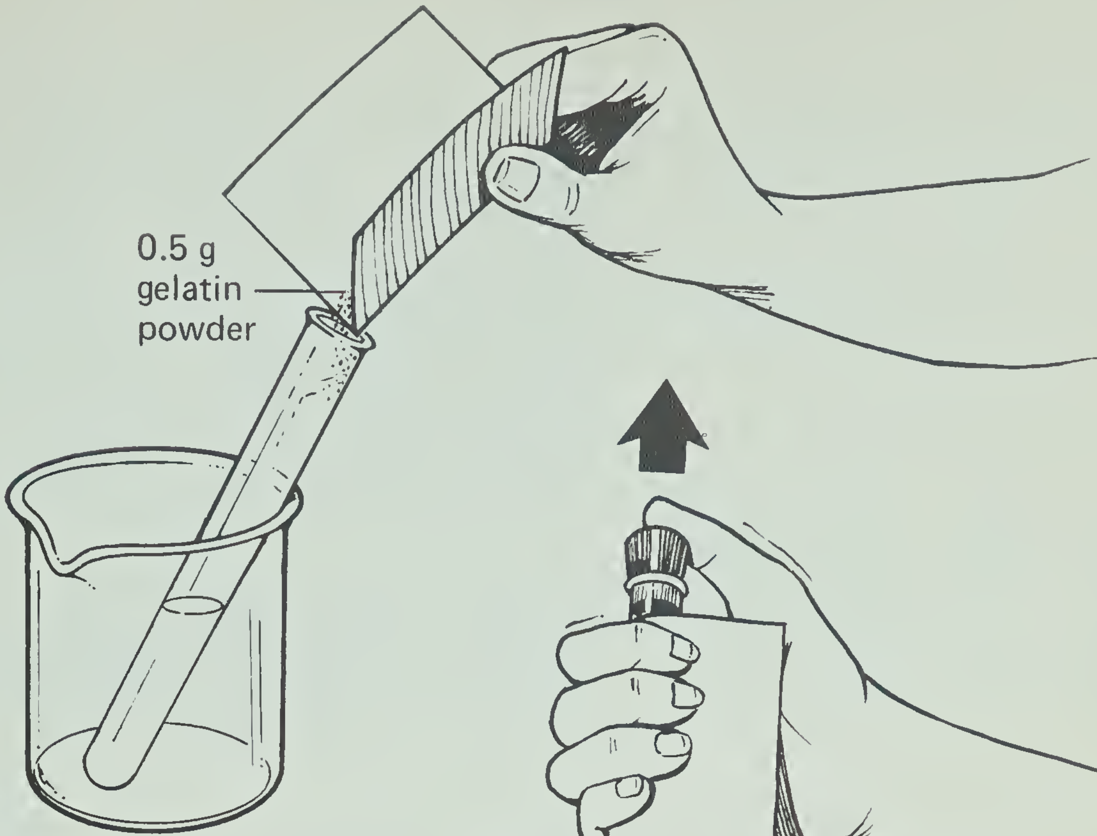
0.5 g of gelatin powder
10-ml graduated cylinder
Bunsen burner or other heat source
medium test tube with stopper
test-tube holder
2 250-ml beakers
paper towel
crushed ice
safety goggles

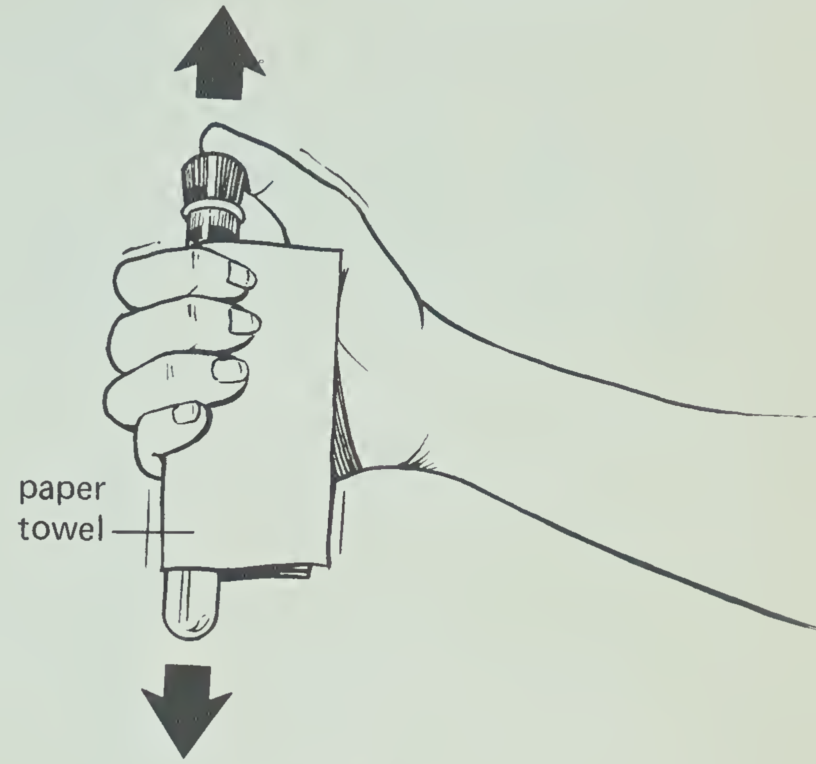
CAUTION	Wear your safety goggles when using the burner. Be careful not to touch a hot test tube with your hands.



A. Put 10 ml of water into the test tube. Using the test-tube holder, heat the water to boiling.

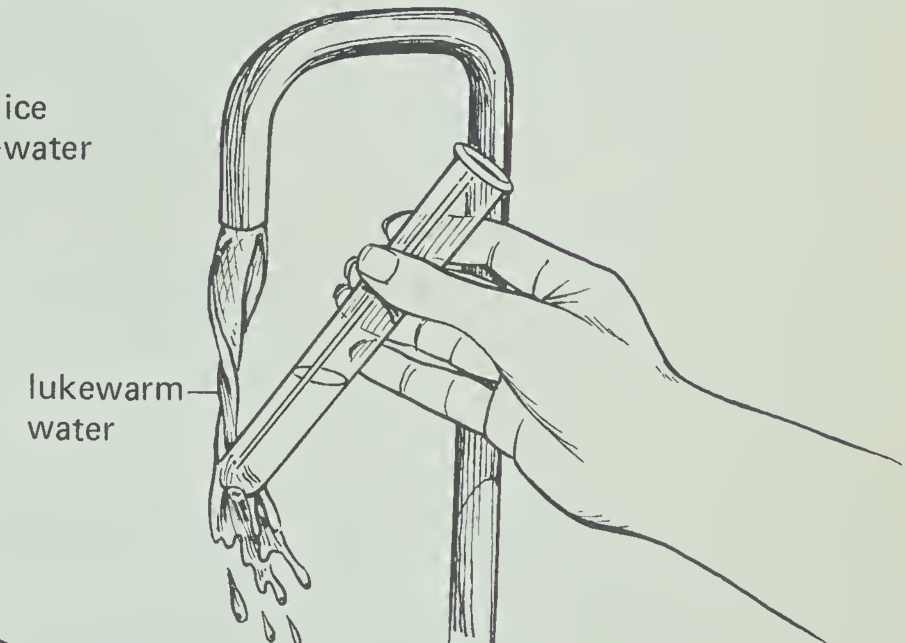
B. Set the test tube with boiling water in a beaker. Pour 0.5 g of gelatin powder into the test tube.



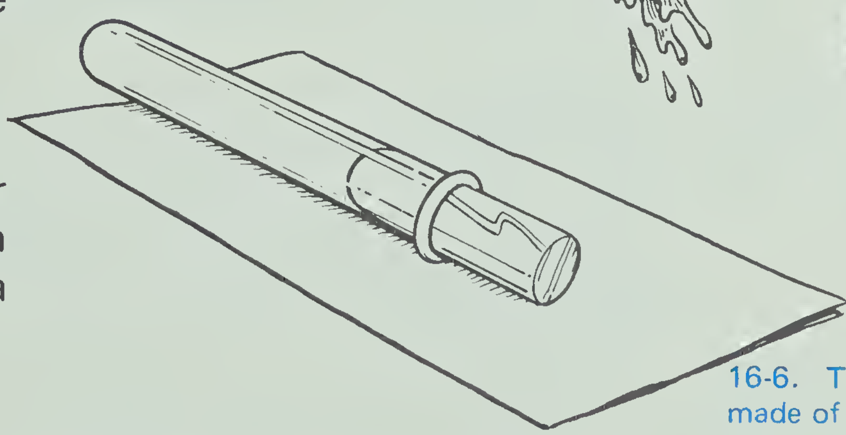


C. Stopper the test tube. Hold the test tube with the paper towel, and shake it until the gelatin powder dissolves.

D. Remove the stopper, and put the test tube into the beaker of ice water.



E. After 3 to 5 minutes, when the liquid has become a gel, remove the test tube from the water.



F. Run some lukewarm water over the test tube. Then slide the gelatin out onto a paper towel. Examine it.

• 16-6. Describe what happens to the protein particles in gelatin granules when water is added.

16-6. The gelatin granules are made of tangled protein molecules with no space between them. When water is added, they swell, and the water works its way between the tangles.

Cream is a suspension of solid butterfat particles in water. When cream becomes slightly sour, you can churn or beat it vigorously. In churning, the butterfat particles get pushed together and get bigger. But tiny droplets of water get caught between the butterfat particles. As you continue to churn, you eventually get water droplets suspended in butterfat. The suspension is reversed — from butterfat particles in water to water droplets suspended in butterfat. It is still a permanent suspension, but now it is *butter*. Butter is a permanent suspension of liquid water in solid butterfat — a solid emulsion.

16-7. Coffee: solid in liquid; salad dressing: liquid in liquid; milk: solid in liquid; butter: liquid in solid; gelatin: solid in liquid; whipped cream: gas in liquid

☆ **16-7. Tell whether each of the permanent suspensions in Figure 16-1 below is a gas-in-liquid, a liquid-in-liquid, a solid-in-liquid, or a liquid-in-solid suspension.**

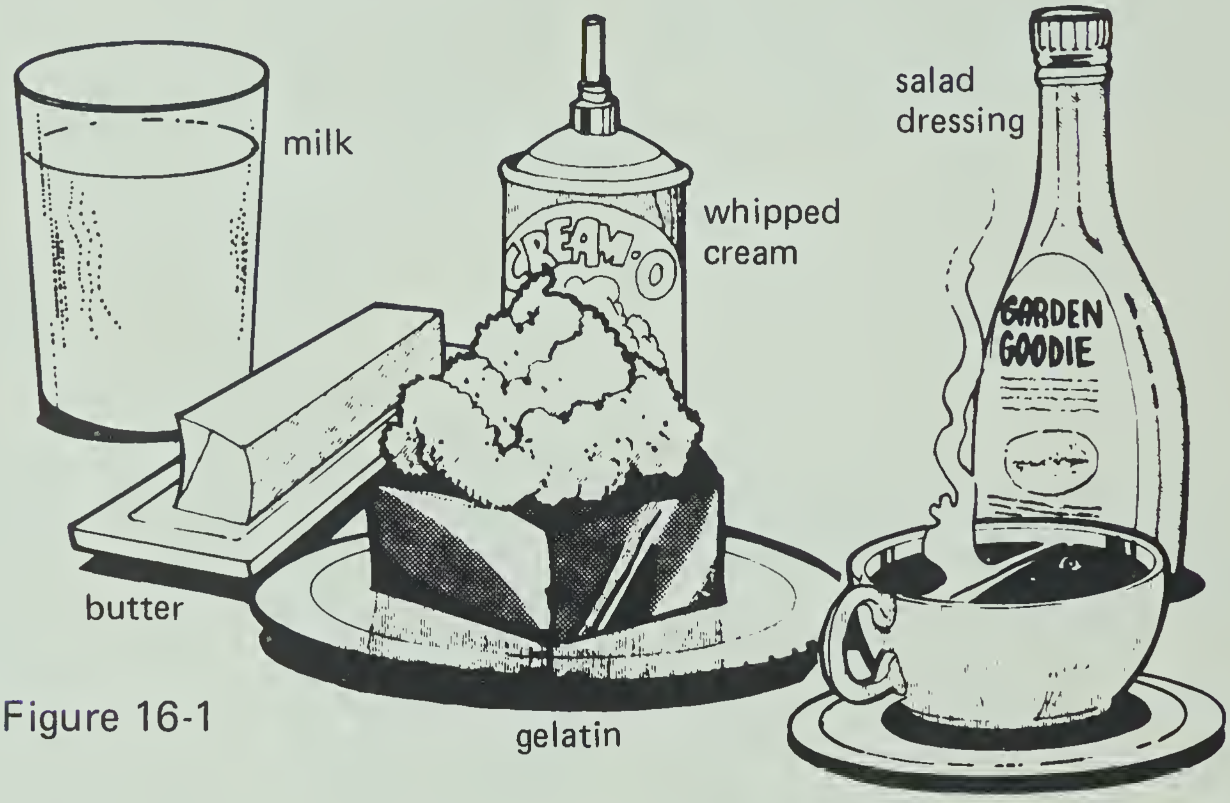


Figure 16-1

ACTIVITY 17: A HOMEMADE INDICATOR

ACTIVITY EMPHASIS: An indicator changes color in acidic and basic solutions.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter.

Activity 2 describes one way to test for acids and bases. Acids turn blue litmus paper to red, and bases turn red litmus paper to blue. The litmus in the paper is called an *acid—base indicator*.

There are many other indicators besides litmus, each with its own colors for acids and bases. Certain flower petals, for example, will change color when you put a drop of acidic or basic solution on them.

Another plant that contains an indicator is red cabbage. To make and use an indicator from red cabbage, you'll need the materials listed on page 75.

red cabbage
distilled water
2 250-ml beakers
6 medium test tubes
sharp knife or shredder
balance
grease pencil
wire sieve or tea strainer
Bunsen burner or other heat source
vinegar in dropping bottle
fruit juice in dropping bottle
sour milk in dropping bottle
household ammonia in dropping bottle
baking soda
cream of tartar
soda straw
ring stand, ring, and gauze
safety goggles

Red cabbage will last about two months if wrapped and refrigerated or about a week if not.

A shredder could be used in place of the knife.

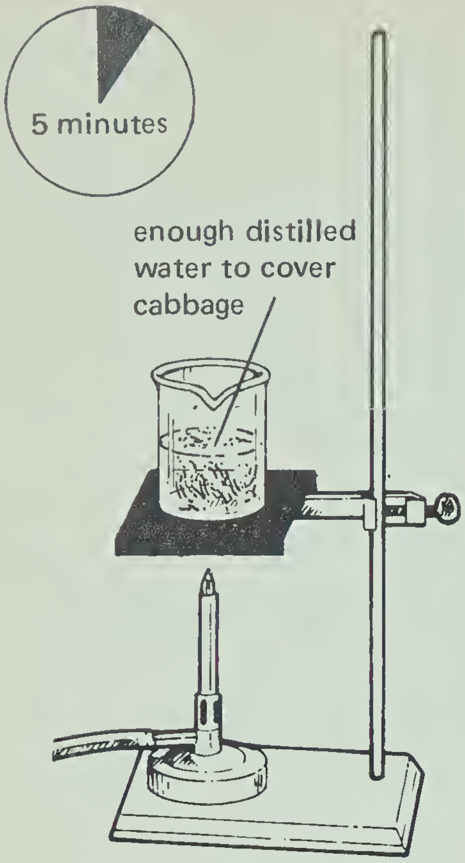
You may want to check "Resource Unit 10: Using Balances" before you begin.

CAUTION	Wear your safety goggles.

A. In your notebook, make a table like this one.

TEST TUBE	SUBSTANCE ADDED	COLOR CHANGE
1	vinegar	
2	fruit juice	
3	sour milk	
4	ammonia water	
5	baking soda	
6	cream of tartar	

B. Weigh out about 40 g of red cabbage. Cut it into small shreds.



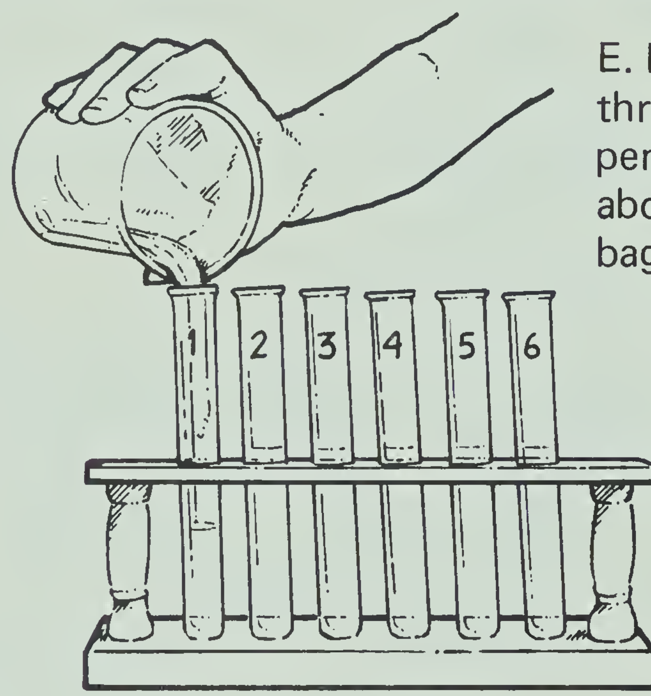
C. Put the shredded cabbage into a 250-ml beaker, and add enough distilled water to cover the cabbage.

D. Heat the cabbage mixture gently for about 5 minutes.

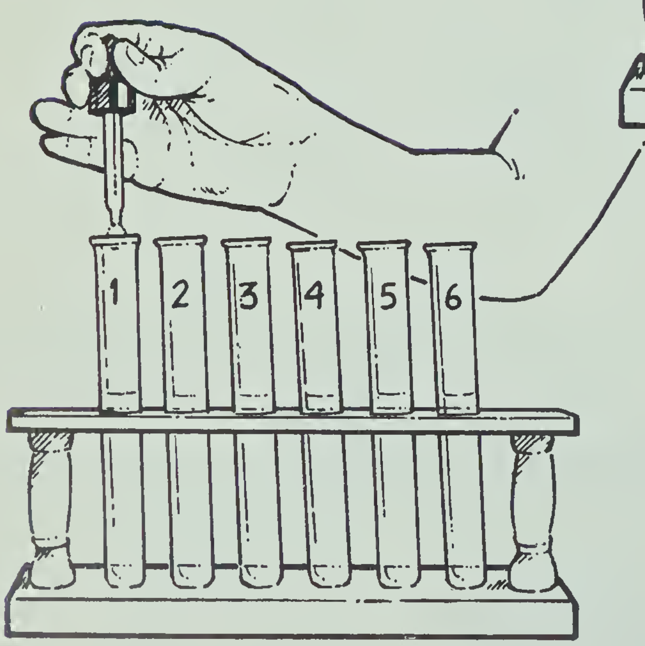
E. Pour the cabbage water through the wire sieve into the other 250-ml beaker. Dispose of the cabbage properly (you won't need it anymore). Save the cabbage water.

17-1. Lavender

● 17-1. What color is the cabbage water?



E. Number the test tubes *1* through *6* with the grease pencil. Fill each test tube about half full of the cabbage water.



F. Add a few drops of each liquid to its test tube (see the table you made in Step A, page 75). Use the straw to add a little of each solid to its test tube. Record the color changes in your table.

17-2. Pink; yellow green

● 17-2. What color did the acidic substances turn the cabbage water? The basic substances?

17-3. It is different colors in acidic and basic solutions.

☆ **17-3. Why is cabbage water useful as an acid–base indicator?**

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